

**EXPERIMENTAL EDITION**

# **SCHOOL CHEMISTRY**

**CLASS 8**

**TEACHER'S GUIDE**

FIRST PUBLISHED

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RESEARCH AND TRAINING

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# SCHOOL CHEMISTRY

## CLASS 8

## TEACHER'S GUIDE

*Prepared by*  
**N C E R T Chemistry Study Groups**



*Under a grant from*  
**THE NATIONAL COUNCIL OF EDUCATIONAL  
RESEARCH AND TRAINING**  
**NEW DELHI**



## PREFACE

The teacher's guide is the companion volume to the text and the laboratory manual published some time ago. The format is essentially the same as the one followed for the teacher's guides of classes 6 and 7. This has been written in that manner because the approach to the study of the subject, as far as the student is concerned, is the same. The major effort in class 8, of course, has been in providing a sort of a connecting link between classes 6 and 7 on the one hand and classes 9 and 10 on the other.

As in classes 6 and 7 the main emphasis has been on students' experimentation and learning through their own observations. The teacher's role is mainly to lay down the general procedure and help in organising the laboratory work. The schedule given for each chapter serves only as a guideline. The teacher has the freedom to organise his teaching schedule, wherever he considers some change is necessary. It is hoped that the background information will be of help to the teacher in presenting the subject in a more facile manner.

The members of the study groups are thankful to the staff of the Department of Science Education, N.I.E., New Delhi, and in particular to Prof. M. C. Pant, for their help in the preparation of this edition. They gratefully acknowledge the help and encouragement given by Prof. D. S. Kothari, the then Chairman of the University Grants Commission, and Chairman, Coordinating Committee, Mr. L. S. Chandrakant, at that time Joint Educational Advisor incharge, NCERT, Mr. Jayaraman, Joint Secretary, Ministry of Education, Government of India, Prof. S. V. G. Iya, Director NCERT. The members of the study groups are thankful to the NCERT for the financial support which has made this project possible.

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# Action of Heat on Substances

## INSTRUCTION MATERIAL

### Introduction

The purpose of chapter I is two fold :

- (i) to acquaint the students with elementary techniques like heating, weighing and manipulation of glass tubing etc.
- (ii) to enable the students to study the action of heat on substances.

### Outline

1. Heat producing no change in composition (1.1)
2. Heat producing change in composition (1.2)

### Concept

Heat produces phase and chemical changes.

### SCHEDULE

Lesson	Experiment	Postlab	Demonstration	Text covered
1	1 & 2		—	—
2	3	Discussion	I	1.1
3	4	Discussion	—	1.2
4	5	Discussion	II	1.2
5	6	Discussion	—	1.2
6	7	Discussion	—	1.2
7	—		—	Review & quiz

## Development

The instruction may start with a demonstration of the various parts of the Bunsen burner and their assembling. During the demonstration, elicit from the students the various sources of heat. If available, it would be worthwhile to demonstrate other types of burners, particularly the fightail burner. Discuss the purpose of each type of burner. The students may then be asked to perform experiment 1.

### Experiment 1 The Bunsen burner

Time required : 30 minutes

Instead of nichrome wire, any other resistance wire may be used.

### Questions and answers

1. Describe the flame obtained when the air inlets are
  - (a) closed and (b) opened
- (a) When the air inlets are closed, the flame becomes sooty due to incomplete combustion of the gas
- (b) When the air inlets are opened, the flame becomes non-luminous due to more or less complete combustion of the gas.
2. Which flame would you use for heating purposes? Why?

The non-luminous flame is used for heating, since it is the hottest flame and does not deposit soot.

Before the commencement of the experiment 2, it is better to demonstrate the manipulation of glass as described in the experiment.

The correct adjustment of the flame, the proper holding of the glass tube, the process of annealing must be clearly brought out. The students may then be asked to perform experiment 2.

### Experiment 2 Working with glass

Time required : 45 minutes

Ask the students to prepare a delivery tube, a tube bent twice at right angles and three or four droppers, which could all be used subsequently.

#### 1.1 Heat producing no change in composition

Ask the students to recall some of the effects of heat on matter, particularly the expansion of matter. The demonstration on phase changes may then be taken up.

## **Demonstration I    Phase changes**

Time required . 30 minutes

### **Materials required**

Beaker (250 ml)	Ice
Thermometer ( $-10^{\circ}$ to $110^{\circ}$ )	
Wire gauze	
Glass rod	
Stop clock	

### **Procedure**

Take about 100 g of crushed ice in a 250 ml beaker and insert the thermometer into it. When the mercury level remains steady note the temperature and record it. Keep the beaker on the wire gauze over a low flame and start the stop-clock. Keep the contents stirred by means of glass rod and record the temperature at intervals of one minute.

Point out to the students that the temperature remains constant till all the ice melts. Introduce the concept of 'phase' (e.g. when the ice is melting, point out the two phases, water and ice). Continue heating and recording the temperature until the water starts boiling. Keep the water boiling for five more minutes. Point out to the students the constancy of temperature at the boiling point.

Draw a graph on the board plotting temperature against time.

Discuss the graph obtained and impress upon students that the temperature remains constant during the phase change of a pure substance. Now the students may perform experiment 3.

### **Experiment 3    Purification of a substance by sublimation**

Time required . 30 minutes

The sample of camphor may be given in packets of 2 to 3 g. If camphor is not available, naphthalene or benzoic acid may be supplied. Recall the process of sublimation and discuss the principle employed in the purification of camphor.

### **Postlab discussion**

Focus the attention of the students on the fact that sublimation is also a phase change. It should be pointed out that in this process there is no change in the composition of the substance.

### **1.2 Heat producing change in composition**

At this stage the students may be introduced through experiment 4, to the action of heat on substances leading to a change in composition.

## Experiment 4    Action of heat on substances

Time required . 30 minutes

0.5 g of each of the materials mentioned in the laboratory manual may be given in separate packets.

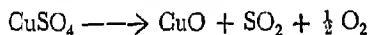
### Preparation of cobalt chloride paper

The cobalt chloride paper may be prepared by soaking filter paper strips in a 5% solution of cobalt chloride and drying them in an oven at 110°. The strips should be preserved in desiccating bottles (bottles containing lumps of fused calcium chloride).

### Preparation of lime water

Lime water may be prepared by shaking 2 to 3 g of slaked lime in 1 litre of water. Let it stand for 2 to 3 hours and then decant. Store the clear solution in a stoppered bottle to protect it from atmospheric carbon dioxide.

Copper (II) sulphate,  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$  (blue in colour) is to be heated very carefully until it becomes almost colourless. Too strong a heating results in the decomposition of copper (II) sulphate into copper (II) oxide (black).



### Postlab discussion

Lead the students to recognise that

- i) substances like zinc oxide do not undergo any change in chemical composition on heating,
- ii) substances like copper (II) sulphate when heated lose water of hydration and undergo some change in composition. But when water is added to the anhydrous substance hydrated copper sulphate (blue) is formed. Thus the process is reversed, and
- iii) substances like lead nitrate and copper (II) carbonate decompose and undergo permanent change in composition.

### Question and answer

Indicate which of the substances you heated

- i) returns to the original state on cooling,
- ii) decomposes but is reformed when the products of decomposition are brought together, and
- iii) seems to have undergone a permanent change.

- i) Zinc oxide returns to the original state on cooling.
- ii) Copper (II) sulphate decomposes, but is reformed when the products of decomposition are brought together
- iii) Copper (II) carbonate and lead nitrate seem to have undergone a permanent change.

Demonstration II can now be taken up to illustrate that heat brings about decomposition. Decomposition is accompanied by loss in weight due to escape of some of the gaseous products formed.

## **Demonstration II     Action of heat on ammonium dichromate**

Time required : 10 minutes

### **Materials required**

Balance	Ammonium dichromate 5 g
Weight box	
Crucible	
Tripod stand	
Clay pipe triangle	
Pair of tongs	
Asbestos sheet	
Sheet of paper	
Desiccator	

### **Procedure**

Place about 3 g of ammonium dichromate in a small crucible and determine the weight correct to a centigram

Insert a 3 cm magnesium fuse into it so that it touches the bottom of the crucible. Place the crucible on an inverted metal tray placed on a large sheet of paper and ignite the magnesium fuse with a Bunsen or other hot flame. A lot of green mass spurts out of the crucible. This demonstration will be very spectacular in a dark room. Collect the green mass thrown out on the paper and transfer it into the crucible. Show that the crucible is almost filled with the product. Allow the crucible to cool in a desiccator.

Weigh the crucible with the product and record the weight. Find the weight of the green mass. Compare it with the mass of the dichromate. Show that it weighs less.

Elicit why there is decrease in the weight even after careful collection of all the green mass. Discuss the reaction and elicit what might have been lost

**Note :** It may be noted that some magnesium oxide is formed and this adds on to the weight. However, it is not considerable. Some chromic oxide formed may be lost due to the vigour of the reaction.

After the above demonstration the correct use of balance has to be demonstrated. Check up if all the balances are in good order before asking the students to proceed with experiment 5.

### Experiment 5 The use of a balance

Time required : 90 minutes

#### Postlab discussion

Ask the students to find the average weight of a peanut. Then plot the average weight of one peanut obtained by each student against his/her roll number. Elicit from the students the answers to the questions given at the end of the experiment in the lab manual.

The students may now proceed to experiment 6

### Experiment 6 Change in weight on heating-I

Time required : 90 minutes

Packets of about 1 g of hydrated barium chloride may be given

If desiccators are not available, desiccating bottles of the type shown in the lab manual can be used. A tightly closed transparent plastic bottle containing fused calcium chloride and a clay pipe triangle bent in the form of a tripod will serve the purpose.

#### Calculation

Wt. of empty crucible	= $W_1$ g
Wt. of crucible + barium chloride before heating	= $W_2$ g
Wt. of crucible + barium chloride after heating	= $W_3$ g
Wt. of barium chloride taken	= $(W_2 - W_1)$ g
Loss of weight in barium chloride	= $(W_2 - W_3)$ g
$(W_2 - W_1)$ g of barium chloride loses $(W_2 - W_3)$ g	

$$\therefore 100 \text{ g of barium chloride losses } \frac{100 (W_2 - W_3)}{(W_2 - W_1)} \text{ g}$$

$$\text{Percentage loss in weight} = \frac{100 (W_2 - W_3)}{(W_2 - W_1)}$$

## Postlab discussion

Ask the students to round off the value of 'x' to the nearest integer since the number of molecules should be a whole number.  
The value of x is 2.

## Question and answer

Using the data and assuming the formula of anhydrous barium chloride to be  $\text{BaCl}_2$ , determine the formula of hydrated barium chloride (At. wt of Ba = 137.41 ; Cl = 35.50)

Formula weight of  $\text{BaCl}_2$  (anhydrous) =  $137.4 + 71 = 208.4$

$(W_3 - W_1)$  g of anhydrous barium chloride combines

with  $(W_2 - W_3)$  g of water

208.4 g i.e., one mole of barium chloride combines with

$$\begin{aligned} \frac{208.4 (W_2 - W_3)}{(W_3 - W_1)} \text{ g of water} \\ = \frac{208.4 \times (W_2 - W_3)}{(W_3 - W_1) \times 18} \text{ moles of water} \\ = x \text{ moles of water (say)} \end{aligned}$$

The formula of hydrated barium chloride is  $\text{BaCl}_2 \cdot x \text{H}_2\text{O}$ .

## Experiment 7 Change in weight on heating-II

Time required : 80 minutes

The magnesium ribbon should be freed of the oxide coating by rubbing with sand paper. Each student should be given about 0.5 g of the metal

### Calculations

Wt. of empty crucible =  $W_1$  g

Wt. of crucible + magnesium =  $W_2$  g

Wt. of crucible + magnesium oxide =  $W_3$  g

Wt. of magnesium taken =  $(W_2 - W_1)$  g

Gain in weight on heating =  $(W_3 - W_2)$  g

$(W_2 - W_1)$  g of magnesium gains  $(W_3 - W_2)$  g

$\therefore$  100 g of magnesium gains  $\frac{100 (W_3 - W_2)}{(W_2 - W_1)}$  g

$$\therefore \text{Percentage gain in weight} = \frac{100 (W_3 - W_2)}{(W_2 - W_1)}$$

### Postlab discussion

Discuss why there is an increase in the weight.

### Question and answer

Using the data, determine the formula of magnesium oxide  
(atomic weight of Mg = 24; O = 16)

$(W_2 - W_1)$  g of magnesium combines with  $(W_3 - W_2)$  g of oxygen.

$\therefore$  24 g (i.e., one g atom) of magnesium combines with

$$\frac{(W_3 - W_2)}{(W_2 - W_1)} \times 24 \text{ g of oxygen}$$

$$= \frac{(W_3 - W_2)}{(W_2 - W_1)} \times \frac{24}{16} \text{ g atom of oxygen}$$

$$= x \text{ g atom (say)}$$

x should be corrected to nearest integer

Formula of magnesium oxide is  $\text{MgO}_x$

The value of x is 1.



## BACKGROUND INFORMATION

### Sources of gas for the Bunsen burner

The usual source of heat in chemical laboratories is the Bunsen burner in which a mixture of air and hydrocarbons obtained by the thermal cracking of kerosene oil or petrol is burnt. Now a days, hydrocarbons obtained as by products in petroleum refineries are available in cylinders—Burshane, Calgas, Esso, Indane, etc, and these may also be used with suitable modifications to the burners

Among the various types of burners used in the laboratory, mention must be made of the fish-tail burner.

For bending glass tubing a fish-tail burner is employed. This gives flat flame (Fig. 1.1a) in which the tube may be evenly heated by gently rotating the tube in it. The same kind of a flat flame may be obtained by fixing a wing top to the burner (Fig. 1.1b).

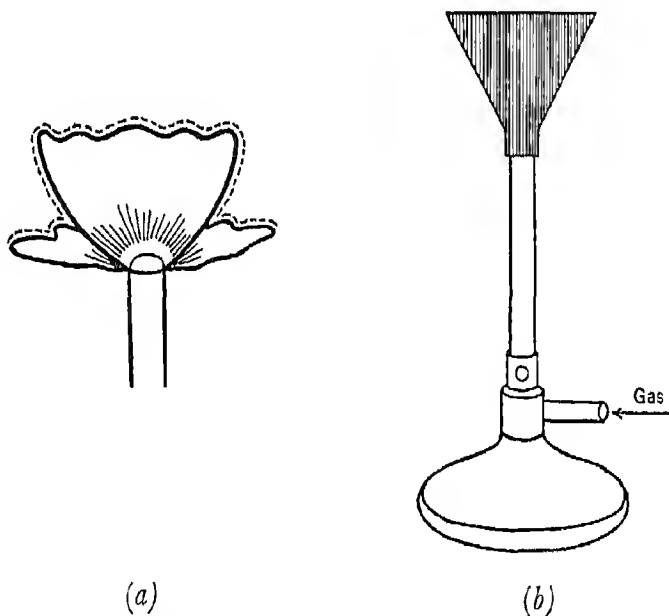
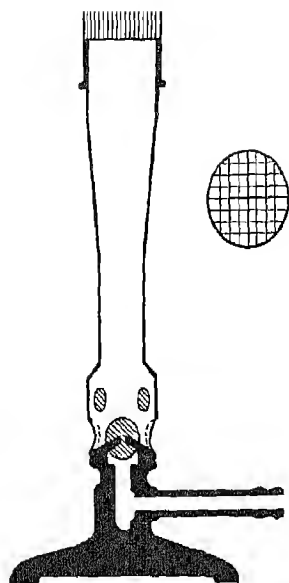


Fig. 1.1 (a) Fish-tail flame (b) Burner with a wing top

The Meker burner (Fig. 1.2) is a larger burner fitted with a wire gauze at its mouth. The air holes are also larger. Each fret of the wire-gauze acts as a tiny Bunsen burner. The Meker burner gives an intense heat over a large area.



## Glass

Glasses are complex silicates of sodium, potassium, calcium and aluminium besides a few other elements. Glasses in which sodium silicates predominate fuse more easily and are known as soda glass or soft glass. Glasses in which potassium silicates predominate are known as potash glass or hard glass and they are more difficult to melt.

Glass is a supercooled liquid. It is not a true crystalline solid. Hence it does not melt sharply on heating but has a wide range of softening temperatures. It is this large temperature range of softening which enables glass to be manipulated easily.

*Fig. 1.2 The Meker burner*

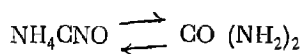
If fused glass is cooled fast, it develops inner cracks or "faults". Hence it is advisable to allow hot glass to cool slowly. This process is called "annealing". This is why after manipulating a glass rod or glass tube, it is covered with a layer of soot so that the glass piece cools slowly.

## 1.2 Heat producing change in composition

When a hydrated crystalline substance is heated, it tends to lose the water of crystallization and become anhydrous. For example,  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$  (blue) on heating loses four molecules of water at  $100^\circ\text{C}$  and forms the monohydrate,  $\text{CuSO}_4 \cdot \text{H}_2\text{O}$  (bluish white). At  $230^\circ$  the fifth molecule of water also is lost and white anhydrous copper (II) sulphate,  $\text{CuSO}_4$ , is formed. On further heating to a temperature of  $650^\circ$  to  $750^\circ\text{C}$ , copper (II) sulphate decomposes into the oxide. It is believed that the four molecules of water lost in the first stage are associated with the cation and other molecule with the anion. Most hydrates lose water on heating. But there are a few exceptions. For example, the hydrate,  $\text{MgCl}_2 \cdot 6\text{H}_2\text{O}$  on heating does not lose water but decomposes giving off hydrogen chloride.

Generally, the water molecules enter the crystal lattice of the hydrated salts. Hence loss of water of crystallisation is followed by the loss of a particular crystalline shape. The tenacity with which the molecules of water are held by the ions in the crystal varies considerably from element to element.

The change of urea to ammonium cyanate is the reversal of the famous experiment of Wohler who obtained urea by the reaction



and synthesised the first organic chemical (urca, which is found in urine) from an inorganic salt. Such changes where there is no change in composition but only a change in molecular structure are known as molecular rearrangements.

Change in the composition of a substance can be brought about by other types of energy besides heat. For example, light energy can cause the decomposition of silver halides and this fact is used in photography. Such reactions are called photochemical reactions. Another example of this type taking place in nature is photosynthesis. Similarly reactions can be brought about by electrical energy as well, - for example, in electrolysis, electroplating, reactions in a cell, reactions under an electric discharge etc. Sometimes the energy released in a certain reaction is enough to initiate another reaction.

## Quiz

Each question given below is followed by four answers. One of them is correct. Tick off (✓) the correct answer and indicate in the bracket provided.

1. An example for phase change is ( )
  - (a) burning of magnesium
  - (b) sublimation of camphor
  - (c) heating of red lead
  - (d) heating of urea.
2. During a phase change the heat supplied is used ( )
  - (a) to overcome the forces of attraction
  - (b) to decrease the thermal agitation of particles
  - (c) to increase the velocity of particles
  - (d) to increase the temperature.
3. During a phase change ( )
  - (a) temperature increases
  - (b) temperature decreases
  - (c) temperature first rises and then falls
  - (d) temperature remains constant.
4. Sublimation takes place when ( )
  - (a) lead nitrate is heated
  - (b) iodine is heated
  - (c) an electric current is passed through acidulated water
  - (d) a zinc rod is placed in silver nitrate solution.
5. An example of a substance which loses weight on heating is ( )
  - (a) zinc oxide
  - (b) litharge
  - (c) copper (II) carbonate
  - (d) sodium chloride.
6. An example of a substance which gains weight on heating in air is ( )
  - (a) magnesium carbonate
  - (b) magnesium
  - (c) platinum
  - (d) red lead,

7. An example of a substance which does not lose water vapour on heating is ( )
- (a) magnesium carbonate (hydrated)
  - (b) red lead
  - (c) copper (II) carbonate
  - (d) cobalt chloride.
8. The substance which evolves oxygen on heating is ( )
- (a) copper (II) carbonate
  - (b) potassium permanganate
  - (c) copper (II) sulphate
  - (d) magnesium carbonate.
9. 84 g of magnesium carbonate gives 44 g of carbon dioxide on heating. The weight of magnesium carbonate required to produce 11 g of carbon dioxide is ( )
- (a) 42 g    (b) 21 g    (c) 12 g
  - (d) 11.5 g
10. When 6 g of magnesium are completely burnt in oxygen we get 10 g of magnesium oxide. The weight of magnesium oxide when 1.5 g of magnesium burns completely in oxygen is ( )
- (a) 2.5 g    (b) 5 g    (c) 7.5 g
  - (d) 0.125 g
11. When 100 g of calcium carbonate is heated 44 g of carbon dioxide is obtained. The number of grams of calcium carbonate required to yield 11 g of carbon dioxide is ( )
- (a) 70 g    (b) 50 g    (c) 25 g
  - (d) 12.5 g
12. 100 g of calcium carbonate on heating yields 56 g of calcium oxide. The number of grams of calcium carbonate to yield 7 g of calcium oxide is ( )
- (a) 3.5    (b) 7    (c) 12.5
  - (d) 28

# Electrical Nature of Matter

## INSTRUCTION MATERIAL

### Introduction

The purpose of this chapter is to establish the presence of matter. An attempt is made to explain the electrical conductivity in solids and solutions.

### Outline

1. Thermionic emission (2.1)
2. Photoelectric emission (2.2)
3. Gas discharge (2.3)
4. Electrical conduction in metallic substances (2.4)
5. Electrical conduction in liquids and solutions (2.5)
6. Relation between quantity of electricity and the amount of material deposited (2.6)

### Concepts

1. Electron is an universal constituent of all forms of matter.
2. Conduction of electricity is due to the flow of electrons and/or of ions.
3. The amount of substance(s) liberated at the electrode is directly proportional to the quantity of electricity passed.

## SCHEDULE

Lesson	Experiment	Postlab	Demonstration	Text covered
1	—	Discussion	I	2.1
2	—	Discussion	II	2.2
3	—	Discussion	III	2.3
4	8 & 9	Discussion	—	2.4 & 2.5
5	—	Discussion	IV	2.5
6	—	Discussion	V	2.6
7	—	—	—	Review & quiz

### Development

#### 2.1 Thermionic emission

The phenomenon of thermionic emission may be illustrated through Demonstration I. The observations made by the students should be discussed in the light of the material given in the text.

#### Demonstration I: Thermionic emission

Time required : 30 minutes

A suitable radio valve with low characteristics (6H6 for instance) which is easily available in radio components shops may be used. A typical assembly required for this demonstration is illustrated in Fig. 2.1

#### Materials required

Radio valve 6H6

Milliammeter (0–300 ma) for filament current

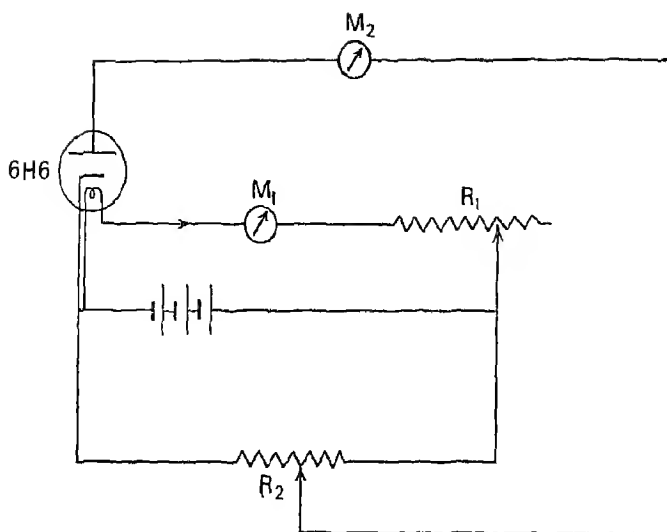
Milliammeter (0–1 ma) for anode current

Wirewound variable rheostat (50–10 watts)

Wirewound variable rheostat (1000–3 watts)

Accumulator 6 volts

Connecting wires with crocodile clips – 6



*Fig. 2.1 Thermionic emission (schematic circuit diagram)*

### Procedure

Assemble the apparatus as shown in the circuit diagram (Fig. 2.1) and keep it ready before the lesson begins. Explain the different components of the apparatus and their functions. The valve 6H6 is a diode consisting of a cathode heated by a hot filament and a plate (anode). The filament is placed very near to the cathode but not touching it. When the filament is connected to an accumulator the current passing through the filament heats it and thus the cathode is heated to a high temperature. The heat supply to the cathode can be adjusted by regulating the current through the filament with the help of the rheostat  $R_1$ . The heated cathode emits electrons which are accelerated towards the plate by voltage applied across the cathode and the plate. Rheostat  $R_2$  controls the voltage across the diode. The milliammeter  $M_2$  measures the electron flow through the diode valve.

Adjust the resistance  $R_2$  about midway between its extreme positions. Increase the current passing through the filament by decreasing the resistance  $R_1$  gradually and note the increase in the reading of the milliammeter  $M_2$  in the plate circuit. The increase of the current strength in the plate circuit indicates the electron flow from the cathode to the anode.

Point out to the students that in this demonstration the number of electrons flowing through the anode circuit depends on the temperature of the filament, the plate voltage being constant.



**Note** The conventional method of showing the current flow in the circuit from positive to negative terminal of the voltage source should be avoided and only the direction of the flow of electrons should be shown.

## 2.2 Photoelectric emission

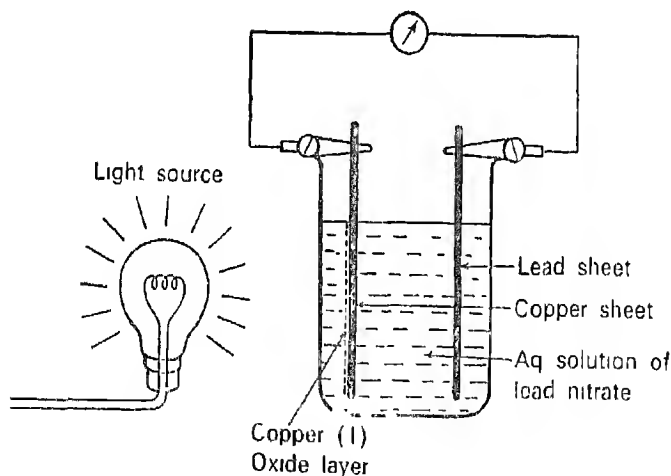
The use of photoelectric cells in various common devices such as the exposure meter, cinematography and television may be mentioned. Metals like cesium, potassium and sodium are found to exhibit photoelectric emission. For the sake of simplicity a liquid photoelectric cell may be demonstrated as described below to illustrate the effect of light on the flow of electrons through an electrolyte.

### Demonstration II: A liquid photoelectric cell

Time required : 30 minutes

#### Materials required

Copper sheet (15 × 3 cm)	0.001 M Nitric acid
Lead sheet (15 × 3 cm)	0.1 M Lead nitrate
Galvanometer	



*Fig. 2.2 Liquid photoelectric cell*

Heat the copper sheet in a Bunsen flame until it is covered with black copper (II) oxide. Then place it in dilute nitric acid for about 5 minutes. Remove it when a red layer is visible. The red layer is copper (I) oxide which is sensitive to light. Immerse the sheet in a jar containing lead nitrate solution. On the other side insert

a clean sheet of lead of the same dimensions and connect the metal strips to a galvanometer. Avoid any direct light on the cell electrodes. The cell should preferably be assembled in semidarkness. Shine a strong light on the copper (I) oxide layer and observe that the galvanometer will show an increase in the current reading. Note what happens to the current when the light is moved away.

It is advisable to prepare the strip with copper (I) oxide layer before the lesson starts.

## 2.3 Gas discharge

The central idea to be brought out in this section is that electron is an universal constituent of all matter. There is no simple demonstration which can show the above fact but it will be better at this stage to introduce to the students the phenomenon of the electric discharge through gases under reduced pressure. It may also help the student in understanding the material given in a later chapter.

### Demonstration III Discharge through gases

Time required : 40 minutes

#### Materials required

A set of cathode ray tubes	}	Tube (1) containing air at atmospheric pressure with electrodes at both ends
		Tube (2) evacuated tube with electrodes at both ends
		Tube (3) evacuated tube with cathode at one end, anode (having a hole in the middle) fitted at the centre of the tube and a fluorescent screen at the other end.

6 volts battery or a suitable D. C. supply

Induction coil or an auto ignition coil

Bar magnet (or a strong horse shoe magnet)

D. C. Ammeter (0-100 mAmp)

Connecting wires, crocodile clips and sand paper

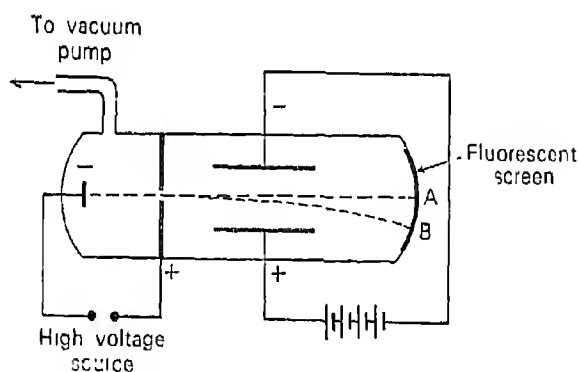
#### Procedure

Connect the secondary terminal of the induction coil to the electrodes of tube 1. No deflection on the ammeter will be observed. Gases at atmospheric pressure do not conduct electricity.

Connect the electrodes of the tube 2 which is evacuated to a low pressure to the induction coil terminals. Ask the students to note the current in the ammeter and to observe the glow between the electrodes.

Use the cathode ray tube with a perforated anode in the centre and a fluorescent screen at the other end (tube 3).

Apply the voltage across the tube and ask the students to note the position of the spot on the fluorescent screen. Explain that the spot is produced by the cathode ray beam. Now bring the north pole of a bar magnet near the cathode ray beam and ask the students to observe the change in the position of the spot on the fluorescent screen. Explain that this is due to deflection of the cathode rays from their original path. Then bring the south pole of the magnet near the cathode ray beam and show that the deflection is in the opposite direction. Take away the magnet, apply the electric field and show that the cathode rays are deflected towards the positive plate (Fig. 2.3). Elicit from the students that the cathode rays are negatively charged.



*Fig. 2.3 Deflection of cathode rays by electric field*

Tell the students that the same results could be obtained using cathodes made of any metal and using any gas in the discharge tube. These results lead to the conclusion that there are particles commonly present in atoms of all elements, and they are negatively charged.

The students should recognise that what they see is the path of the electrons and not the electrons themselves. The glow on the glass wall of the discharge tube is produced by the bombardment of the wall by electrons.

After the demonstration, develop the matter along the lines of the text. Very little historical background need be given here. Following points should be brought in the discussion :—

- 1) The electron has a very small mass,  $1/1837$  of that of a hydrogen atom. Mass of hydrogen atom =  $1.6 \times 10^{-24}$  g. Mass of electron =  $9.1 \times 10^{-28}$  g.
- 2) It carries the smallest known quantity of negative electric charge. This is the unit of electricity ( $1.6 \times 10^{-19}$  coulomb or  $4.8 \times 10^{-10}$  electrostatic units).
- 3) Electrons are the common constituents of the atoms of all elements. We can cause emission of electrons from atoms in different ways.
  - i) as above by an electric discharge through gases,
  - ii) by heating metals to red heat and
  - iii) by the action of light on certain metals,

Electrons obtained from any element by any method are identical in mass and charge. They may however have a different initial velocity. The electron is one of the fundamental particles of matter.

## 2.4 Electrical conduction in metallic substances

The development of this section should start with experiment 8 on electrical conduction through solids.

In a brief prelab discussion demonstrate the assembling and use of the apparatus. The students should be asked to take precaution against damaging of the cells due to short circuiting.

### Experiment 8 : Electrical conductivity of solids

Time required : 25 minutes

Sufficiently large crystals ( $\approx 1$  cm long) of copper sulphate and alum should be supplied so that the electrode terminals are not short-circuited during the experiment. The teacher may supply additional samples available in the school for observing their electrical conductivity. A few substances which may be easily available and which are suitable for this experiment are mentioned here : coins, pieces of marble, sulphur, wax, plastic sheet, candy drops, lumps of iron sulphide, lumps of rocksalt, wooden chips, charcoal, graphite etc.

## Postlab discussion

Write on the blackboard students' observations and see whether they have classified the materials correctly. The discussion should not be limited to the classification. Explain why metals are good conductors and plastics, cardboard etc are non-conductors. Point out that metals possess mobile electrons and conduction is due to movement of these electrons.

When a potential is supplied across the ends of a conductor the mobile electrons get displaced and move along the circuit. Thus a flow of electrons results. During this flow a certain number of electrons enter the metal from the negative terminal and the same number of electrons leave the metal and enter the positive terminal. Thus during the conduction of electricity the electrons are not removed from the metal but are maintained in a continuous motion i.e kept moving in a circuit. There is no visible transfer of matter, or any change in the nature of the conducting substance, during metallic conduction. On the other hand substances like plastics, sulphur, wax, rubber do not contain mobile electrons and so they are non-conductors.

## Question and answer

Based on your observations, classify the substances into conductors and non-conductors of electricity,

Conductors	Non-conductors
Copper	Cardboard
Iron	Glass rod
Lead	Copper sulphate crystal
Wire with both ends bare	Alum crystal
	Wire with plastic insulation at both ends
	Wire with one end bare

## 2.5 Electrical conduction in liquids and solutions

The teacher should start this section with experiment 9 followed by demonstration IV.

## Experiment 9 : Electrical conductivity of liquids and solutions

Time required : 30 minutes

The following solutions should be supplied to the students.

- a) Benzene, alcohol and kerosene (50 ml of each liquid may be given).
- b) Aqueous solutions of
  - i) 0.2 M hydrochloric acid (16 ml concentrated acid per litre)
  - ii) 0.2 M sodium hydroxide (8 g solid dissolved in 1 litre)
  - iii) 0.2 M potassium iodide (33 g solid dissolved in 1 litre)
  - iv) 0.2 M sodium chloride (12 g solid dissolved in 1 litre)
  - v) 2% sugar solution (20 g solid dissolved in 1 litre)
- c) Solution of hydrogen chloride in benzene is prepared by bubbling the dry gas through benzene for about 10–15 minutes.

Hydrogen chloride gas is prepared as follows.

About 100 g of common salt are taken in a flat bottomed flask and concentrated sulphuric acid is dropped slowly through the thistle funnel. The gas generated is dried by passing it over calcium chloride kept in a U-tube. The dried gas is bubbled through benzene.

**Note :** The reaction between sodium chloride and sulphuric acid proceeds in the cold. However, if a more vigorous supply of hydrogen chloride gas is required, the mixture may be heated gently.

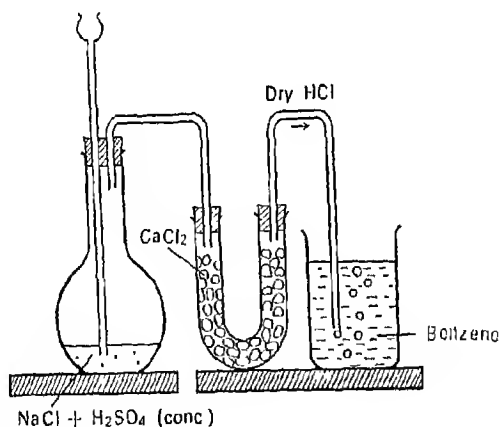
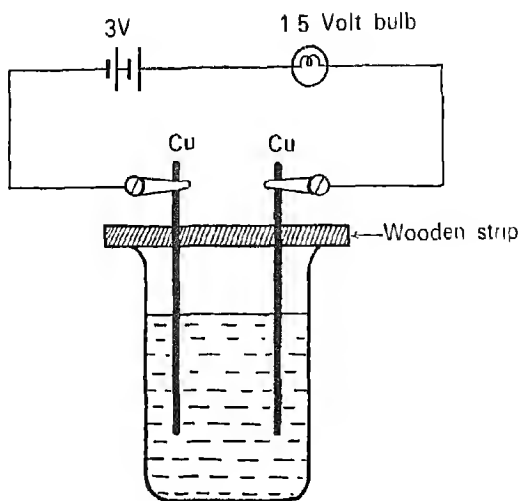


Fig. 2.4 Dissolution of hydrogen chloride gas in benzene

The apparatus indicated in Fig. 13, page 23 of the laboratory manual may be used. It is necessary to see that the electrode terminals do not get short circuited during the experiment. An improved form of apparatus which eliminates short-circuiting is given below (Fig. 2.5).

Copper electrodes may be fixed about 2-3 cm apart from each other in a wooden strip. This type of electrode-holder should be placed on the beaker containing the liquid.



*Fig. 2.5 Electrical conductivity of liquids and solutions*

A brief prelab discussion may be necessary to explain and demonstrate the assembling and working of the apparatus, and precautions to be taken against short-circuiting. Use a torch bulb of 1.5 volts in place of a 3 volt bulb.

### **Postlab discussion**

It should not be limited to the classification of liquids. The following points may be brought out from the students' observations -

a) All solutions do not necessarily conduct electricity. Conductivity is dependent upon the free movement of ions.

b) Dissolving a solid or a gas in a solvent does not always give a solution containing ions.

## Question and answer

Based on your observations classify the substances into conductors and non-conductors of electricity.

### Conductors

Aqueous solutions of  
hydrochloric acid,  
sodium hydroxide,  
potassium iodide,  
Common salt.

### Non-conductors

Distilled water  
Benzene  
Kerosene  
Alcohol  
Sugar solution  
Solution of hydrogen chloride in benzene.

**Note :** Alcohol may show very feeble conductivity if it contains considerable amount of water. Distilled water shows feeble conductivity.

## Electrical conduction through fused salts

In demonstration IV, in the place of lead bromide any salt such as  $\text{SnCl}_2$  ( $243^\circ\text{C}$ ),  $\text{PbI}_2$  ( $412^\circ\text{C}$ ) or  $\text{NaNO}_3$  ( $316^\circ\text{C}$ ) which melt at a reasonably low temperature could be used.

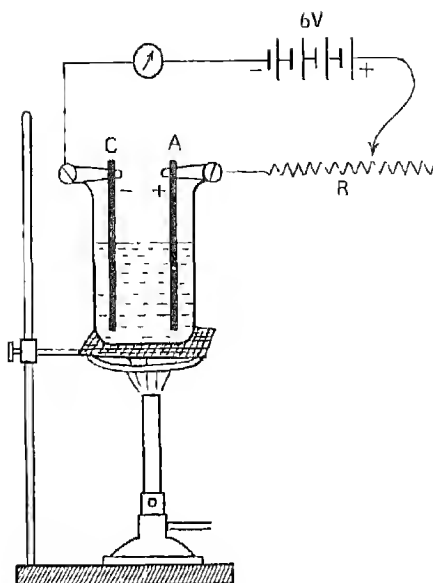
## Demonstration IV · Electrical conduction through a fused salt

Time required : 30 minutes

### Materials required

6 volt battery	Lead bromide
D. C. Ammeter (0-0.2 Amp)	
Rheostat	
Carbon rods (from spent dry cells)	
Clips to hold electrodes	
Beaker (100 ml)	
Iron stand with clamp	
Bunsen burner	
Wire gauze	





*Fig. 2.6 Electrical conductivity through a fused salt*

### Procedure

Take the salt in a beaker. Place the electrodes as shown in the Fig. 2.6. Gently heat the beaker on a Bunsen burner until the salt melts.

Allow the contents to cool and solidify. Connect the electrodes to the battery as shown in the diagram. Let the students observe that solid lead bromide does not conduct electricity. Heat the beaker again till the salt melts. Show to the students that in the molten state the salt conducts electricity.

**Note 1 :** Avoid overheating since the salt may decompose and give out toxic fumes.

**Note 2 :** Students may be inquisitive to know why salt was melted initially. Tell them that this was done to drive away air (bad conductor) between lead bromide particles and to obtain a lump of the salt in contact with the electrodes.

Discuss the results with the students. Elicit from them the following facts: Metals are good conductors. Ionic substances like lead bromide do not conduct electricity in the solid state. But molten lead bromide as well as an aqueous solution of lead bromide are conductors. This means that they contain ions which are mobile and so are free to carry current. In solid lead bromide, the ions are present but they are not free to move and carry the current.

This method of conduction of electricity in aqueous solution and melts is 'ionic' or 'electrolytic'. At this stage a comparison may be made between 'ionic' conduction through liquids and the 'electronic' conduction through metals. The principal idea to be emphasized in the discussions on section 2.5 is that an electric current is a flow of electrons or ions.

## 2.6 Relation between quantity of electricity and the amount of material deposited.

Since electricity is carried by ions, (during the electrolysis of ionic solids in fused state and their aqueous solutions) whose charges are simple integral multiples of unit positive and negative charges depending upon their nature, it is reasonable to expect a quantitative relationship between the quantity of electricity passed and the amount of products formed. The relationship can be brought out through demonstration V given below.

It is advisable to start this demonstration before the school hours, so that the results will be available after 1 period.

### Demonstration V: Relation between the quantity of electricity and the masses of products

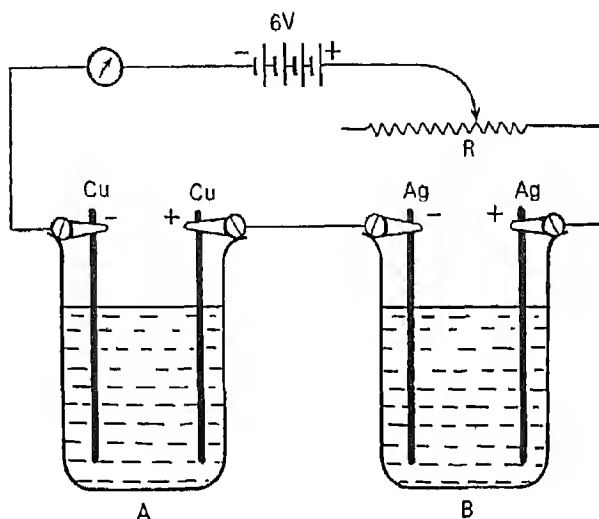
Time required : 30 minutes

#### Materials required

6 volt battery	0.05M copper (II) sulphate solution (1.25g salt dissolved in 100 ml water)
Rheostat - ( 5 $\Omega$ resistance )	
D.C. Ammeter ( 0 - 0.5 amps )	
Beakers (100 ml ) 2	0.05 M silver nitrate solution (0.8 g salt in 100 ml)
Connecting wires	Ethanol 50 ml
Copper foil electrodes 2 (5×3 cm with electrode-holder)	Acetone 50 ml
Silver foil electrodes 2 (5×3 cm with electrode-holder)	
Stop watch	

## Procedure

Clean the electrodes, weigh them accurately and set up the apparatus as shown in Fig. 2.7. The beaker A contains 0.05 M copper (II) sulphate solution and copper electrodes, and beaker B contains 0.05 M (approx) silver nitrate solution and silver electrodes. Start the current and the stop watch. Using the rheostat R adjust the



*Fig. 2.7 Relation between the quantity of electricity and masses of products*

current density at the cathode to about 10 milliamp.  $\text{cm}^{-2}$ . (total current in milliamps divided by area of cathode in  $\text{cm}^2$ ). For this, with electrodes given above, about 0.1 amp current will be required as the area dipping in the solution will be about  $10 \text{ cm}^2$ . Adjust the current strength to 0.1 amp with the help of the rheostat.

After one hour remove the cathodes, wash them by dipping first in distilled water and then in acetone. Since the silver deposit may not adhere well to the cathode, care has to be taken so that no silver is lost during washing. Find the increase in the weight of both the cathodes and compare the ratio of the gain in weights to the ratio of atomic weights of the metals. Calculate the quantity of electricity passed through the circuit from the current strength and time. From these results calculate the quantity of electricity required to deposit one mole of copper and one mole, of silver.

It follows that there must be a direct relation between the number of atoms (or gram atoms or moles of substances) deposited (or liberated) at the electrodes and the number of electrons in the circuit. In other words, the mass of the given substance deposited at the electrodes will be proportional to the quantity of electricity,

$$Q = n.e.t.$$

where  $n$  = number of electrons passing a reference point in the circuit per second,

$e$  = charge of electron ( $1.6 \times 10^{-19}$  coulomb) and

$t$  = time in seconds.

The students can thus associate a finite amount of electricity with the liberation of one mole of any element during electrolysis. In a simple case when an atom is formed from the ion by the gain or loss of one electron, the quantity of electricity required to liberate one mole of the element ( $6.02 \times 10^{23}$  atoms) will be the charge carried by one mole of electrons (i.e.  $6.02 \times 10^{23}$  electrons). This quantity of electricity is known as one Faraday of electricity.

$$F = 6.02 \times 10^{23} \times 1.6 \times 10^{-19} \text{ i.e. } 96500 \text{ coulombs.}$$

The numerical problems given in the text may be discussed in the class to make these ideas more clear.

## BACKGROUND INFORMATION

### 2.1 Thermionic emission

Thermionic emission takes place when heat is supplied to metal atoms. The minimum quantity of energy required to remove an electron from the metal surface depends upon the nature of the metal, its surface condition, absorbed films and metal impurities. The phenomenon of thermionic emission was used by Sir John Fleming in a device used to rectify electric current (convert A.C. to D.C.). Such an arrangement is called vacuum tube rectifier or vacuum diode (arrangement with two electrodes). A vacuum diode consists of an anode and a cathode enclosed in high vacuum. The cathode is heated to the required temperature either by passing an electric current (direct heating) or by placing it near a hot filament (indirect heating). The heated cathode emits the electrons which are collected by the anode under the voltage applied across the two electrodes. In a directly heated cathode a tungsten or thoriated tungsten wire is heated to about  $2000^{\circ}$  or  $1300^{\circ}\text{C}$ . respectively, by passing a direct current. For the indirectly heated cathode, a nickel foil covered with a film of alkaline earth oxides is heated to  $700\text{--}900^{\circ}\text{C}$ . by a filament close to but not touching it. The filament is made of a resistance wire and is heated by the passage of either an alternating or direct current. The number of electrons emitted by the cathode depends on the work function of the metal and its temperature. The anode attracts the electrons emitted by the cathode and thus a current flows through the diode. The strength of the current depends upon the surface area of the cathode, the area and nature of the anode surface and the applied voltage between the electrodes.

### 2.2 Photoelectric emission

Photoelectric emission differs from thermionic emission only in that the energy required for the emission of electrons is supplied by light (of suitable wavelength). When the energy associated with the incident light photon is greater than that required to release an electron (threshold energy known also as the work function) the electron is released from the surface with a kinetic energy equal to the excess energy supplied. ( $= \text{energy of photon} - \text{threshold energy}$ ). Thus, if  $E$  is the incident energy and  $E'$  is the threshold energy, the kinetic energy of the released electron is  $(E - E')$ . The number of emitted electrons depends on the surface area illuminated and the intensity of light.

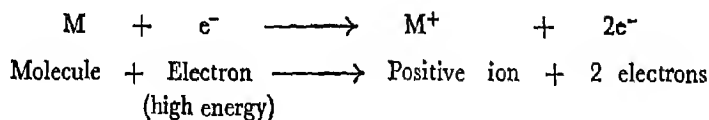
Metals such as cesium and rubidium are generally used in photoelectric devices.

## 2.3. Gas discharge

### Mechanism of conduction of electricity through gases

At ordinary pressure gases consist mostly of molecules which are electrically neutral. Of these millions of molecules an extremely small fraction exist as positive ions and free electrons. Hence gases are ordinarily non-conductors even when high voltages are applied. The extremely small number of electrons and an equal number of positive ions present in a gas are moving about in all directions in a random way like all the molecules. In this process the electrons collide with the neutral molecules frequently. On application of a high voltage across the electrodes the random motion is partially changed into a regular movement. The electrons move towards the anode and the positive ions towards the cathode. In this process the electrons in particular get accelerated (i.e. they acquire higher and higher speeds). This is similar to the increase in the speed which a piece of stone gains while falling to the earth due to gravity. At high pressures the number of molecules in a given volume is large and the electron has to undergo collisions with a large number of molecules before it travels long. At each collision it loses a part of its energy to molecules with which it collides. But as the pressure is lowered, the number of molecules in a given volume diminishes and hence the number of electron-molecule collisions is greatly reduced. This permits electrons to travel a longer distance before colliding. Hence the electrons can get accelerated to a higher energy.

As the pressure in the discharge tube is further reduced a stage is reached when the electrons gain sufficient energy to knock out an electron from a molecule during collision as shown below.



This process repeats and the number of electrons and positive ions multiplies. In this way with the increased number of free electrons and positive ions the gas begins to conduct, that is, an electric discharge takes place. Because of their higher mass positive ions move very slowly compared with the electrons. Therefore in a gas discharge the electricity is conducted mainly by the electrons.

## 2.4 Electrical conduction in metallic substances

Different metals and alloys have different electrical conductivities. Metals like platinum, silver, copper and aluminium are some of the best conductors. Alloys like nichrome are poor conductors and for the same applied voltage they get heated more than nickel and chromium. Plastics, mica, rubber and porcelain are non-conductors and hence are used as insulating materials.

Metals conduct electricity by the flow of electrons. In metals, some valence electrons are relatively free to move. These mobile electrons are effective in holding the crystal together because wherever they move, they are simultaneously attracted by two or more nuclei. In substances like wax, plastics, rubber etc the electrons are localised (i.e. they are not free to move) in space between a particular pair of bonded atoms. Therefore covalent substances do not normally conduct electricity.

The behaviour of carbon is a special one. One allotropic form of carbon namely 'diamond' is a non-conductor like most other non-metals. On the other hand, its other allotrope 'graphite' conducts electricity

In diamond (Fig. 2.8 A) each carbon atom is at the centre of a regular tetrahedron. It is linked with four other carbon atoms present at the corners of a tetrahedron. Each of these carbon atoms would appear to be at the centre of another tetrahedron. Thus the structure is extended in three dimensions and the entire diamond crystal is to be considered as a giant molecule. In this symmetrical structure the carbon-carbon bond distance is  $1.54 \text{ \AA}$ .

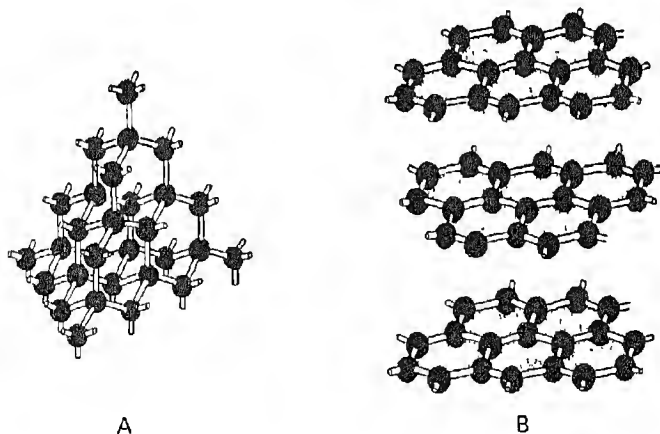


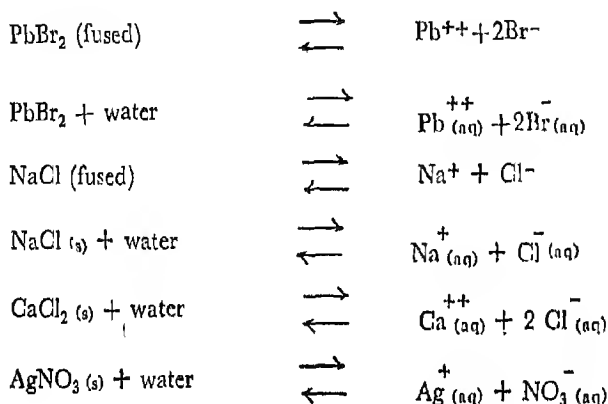
Fig. 2.8 Structure of (A) diamond and (B) graphite

All the four valence electrons of every carbon atom in diamond are symmetrically used for forming covalent bonds of the same strength. The electrons are strongly localised between adjacent carbon atoms and there are no free electrons to conduct electricity.

Graphite (Fig. 2.8 B) is a two dimensional laminar network of atoms in which only three electrons of a carbon atom are used symmetrically in forming bonds with three other carbon atoms in a plane. The planes are arranged one above the other. The adjacent planes are held together by van der Waals' forces. The fourth electron of each carbon atom is relatively delocalised (free to move). It accounts for the electrical conduction through solid graphite.

## 2.5 Electrical conduction in liquids and solutions

When an ionic solid is fused or dissolved in water, it dissociates into positive and negative ions as shown below :



In aqueous solution, the ions are hydrated due to the polar nature of the solvent (water).

The ions thus dissociated recombine to a certain extent to form clusters, which in turn dissociate again into ions. The state of equilibrium may be shown by the arrows pointing in both directions.

Water is a very poor conductor of electricity. Yet when an ionic solid dissolves in water, the solution conducts readily. This is due to formation of ions in the solution. Conduction is determined by the number of ions per unit volume of the solution as well as their mobilities.

When a current is passed through an aqueous solution of an electrolyte different reactions take place at the electrodes and some products are liberated. In general, the reactions at the cathode produce hydrogen or a metallic element. The anode reactions involve liberation of oxygen or any other non-metallic element or dissolution of the anode. In aqueous solutions, the liberation of hydrogen is always accompanied by an increase in pH near the cathode and liberation of oxygen by a decrease in pH near the anode.

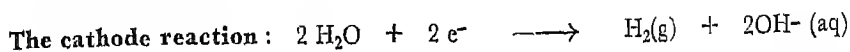
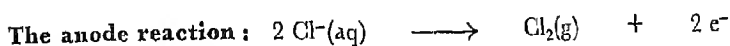


It is known that water conducts electricity to a very slight extent because of the dissociation of water into  $H^+$  and  $OH^-$  ions.



( One molecule out of about  $10^7$  molecules dissociates )

The presence of traces of impurities also alters the conductivity of water. The mechanism of conduction by aqueous solutions of electrolytes becomes more complicated in the case of dilute aqueous solutions due to the participation of  $H^+$  and  $OH^-$  ions of water. For example, in the electrolysis of a dilute solution of sodium chloride we get chlorine gas at the anode and the hydrogen gas at the cathode.



## 2.6 Relation between quantity of electricity and the amount of material deposited

The chemical change brought about in electrolysis by the transfer of electric charge through ions is successfully interpreted in terms of the atomic theory. In fact the stoichiometric relations observed in electrolysis have provided chemical evidence to the atomic theory and the proof for the existence of the electron. When Michael Faraday first performed his experiments on the electrolysis in the early nineteenth century, the atomic theory had already been proposed, but the existence of the electron was not known. There was no reason to suspect that electricity consisted of discrete particles. Faraday observed that the quantity of electricity required to deposit a given weight of an element from solutions of its different compounds was always equal to a constant, or some simple multiple of this constant (Faraday's first law of electrolysis). Restated in terms of the atomic theory, it means that the number of atoms of an element deposited by a fixed quantity of electricity is constant or a simple multiple of this constant. This suggests that there are 'quanta' and 'packets' of electricity which are transferred during electrolysis; one, two or three per atom deposited, but never a fraction of this quantum of electricity which was later identified as the charge on the electron.

Faraday's second law of electrolysis is based on the observation that the ratio of the number of moles of different elements deposited by the same amount of electricity is a simple whole number. This result shows that a certain amount of electricity will deposit a fixed number of atoms, or some simple multiples of this number, whatever the element is. Both of these observations by Faraday lead us to conclude that the atom of any element can carry only an integral multiple of a fixed quantity of electricity. The fixed quantity of electricity is the same as the charge of an electron present in any atom.

## Quiz

Each question given below is followed by different answers. Only one of them is correct. Select the correct answer and indicate in the brackets provided.

1. The phenomenon of emission of electrons from lithium when light of suitable energy falls on it is called ( )
  - (a) gas discharge
  - (b) photoelectric emission
  - (c) thermionic emission
  - (d) electrolytic conduction.
2. The phenomenon of releasing negatively charged particles from a hot cathode is known as ( )
  - (a) electronic conduction
  - (b) thermionic emission
  - (c) electrolytic conduction
  - (d) photoelectric emission.
3. Passage of electricity through metals by the flow of electrons is known as ( )
  - (a) thermionic emission
  - (b) gas discharge
  - (c) electronic conduction
  - (d) electrolytic conduction.
4. An example of a non electrolyte is ( )
  - (a) sodium chloride
  - (b) lead bromide
  - (c) benzene
  - (d) copper (II) chloride.
5. During electrolysis the quantity of electricity required to deposit one mole of silver is ( )
  - (a) one faraday
  - (b) two faradays
  - (c) three faradays
  - (d) four faradays.



12. Electrical conduction in metals is thought to be (            )
- (a) the same as that in solid electrolyte
  - (b) the same as that in molten electrolyte
  - (c) dependent on the movement of ions
  - (d) dependent on the movement of valence electrons.
13. Write equations for the reactions that take place when molten tin (II) chloride is electrolysed.
14. Molten sodium chloride is electrolysed and 0.115 g of sodium is deposited at the cathode. How many grams of chlorine will be formed at the anode? How many coulombs of electric charge is passed through the solution? Write the anode and cathode reactions.
15. Some concentrated sodium chloride solution is electrolysed using a current of 0.50 amperes. How much chlorine will be obtained when the current has been flowing for one hour?
16. What weight of silver would be deposited by a current of 15 amperes flowing through silver nitrate solution for one hour?
17. How many coulombs of electricity are required to deposit 100 g of copper from copper (II) sulphate solution?
18. What mass of chlorine is produced by passing a current of 0.25 amperes for 6 hours through an aqueous solution of sodium chloride?
19. What weights of sodium and chlorine will be formed when 10,000 coulombs of electricity is passed through molten sodium chloride?
20. A current of 0.25 amp, is passed through an aqueous solution of sodium chloride for 3 hours. Calculate the volume of chlorine (at S.T.P.) produced. The experiment is conducted at 27°C and a pressure of 1 atmosphere.

# Structure of Atom

## INSTRUCTION MATERIAL

### Introduction

In the preceding chapter we learnt that electrons are common constituents of all atoms. In this chapter we will study some other subatomic particles and their arrangement in an atom. The model of the atom developed here will be used subsequently to interpret the chemical behaviour of atoms

### Outline

1. The positive ions (Positive rays) (3.1)
2. Radioactivity (3.2)
3. The atomic nucleus (3.3)
4. Atomic number (3.4)
5. The neutron (3.5)
6. Isotopes (3.6)
7. Arrangement of electrons around the nucleus (3.7)

### Concepts

1. Atom is not the ultimate particle of matter but it composed of smaller particles like electrons, protons and neutrons.
2. Radioactivity is the spontaneous emission of  $\alpha$  or  $\beta$  and  $\gamma$  rays by atoms of some elements.
3. The atomic nucleus is a small region in an atom in which nearly the whole mass and all the positive charges of the atom are situated,
4. Electrons are arranged in different shells surrounding the nucleus.

5. Atomic number gives the number of protons and also the number of electrons present in an atom.
6. The atomic number characterises the chemical identity of an element
7. Isotopes are atoms of the same element having the same atomic number but different mass numbers.

### SCHEDULE

Lesson	Experiment	Postlab	Demonstration	Text covered
1	—	—	—	3.1, 3.2
2	—	—	—	3.3, 3.4
3	—	—	—	3.5, 3.6
4	—	—	—	3.7
5	—	—	—	Review and quiz

### Development

#### 3.1 The positive ions (Positive rays)

Recall the discussion on cathode rays (Chapter 2) to establish that electrons are present in all atoms. Indicate the possibility of atom having positively charged particles as a logical consequence of electrical neutrality of atoms,

Describe Goldstein's experiment which revealed the existence of positively charged particles in the discharge tube. Introduce the idea of proton which is the smallest positively charged particle obtained when hydrogen gas is used in the discharge tube. It stands for unit positive charge and has a mass nearly equal to the mass of a hydrogen atom, which is 1837 times that of an electron. Emphasise that positively charged particles obtained when other gases are used in the discharge tube are not protons.

#### 3.2 Radioactivity

Further evidence for the existence of electrically charged particles, positive as well as negative, in an atom can be provided through the study of radioactivity. The discussion should be brief since it is not intended here to give a detailed information about radioactivity. Point out that certain atoms spontaneously emit radiations of different types. Some of them contain positively charged  $\alpha$  - particles (helium nuclei) and some of them, negatively charged  $\beta$ -particles (electrons) and these  $\alpha$  and  $\beta$  particles are accompanied by high energy, short wave length  $\gamma$  - radiations. Point out that this

phenomenon is shown generally by elements such as radium, uranium, thorium. This indicates that the nuclei of the atoms of these elements are unstable.

### 3.3 The atomic nucleus

Discuss then how these different subatomic particles—electrons, protons and neutrons are arranged in an atom. Present Thomson's model in brief (see background information, page 44). Describe the experiment of Rutherford and point out how Thomson's model is untenable. Describe how Rutherford logically deduced the concept of nucleus. The nucleus consists of protons and neutrons. The relative sizes of the atom and the nucleus may be indicated with suitable analogies like a mosquito at the centre of a big auditorium. The former represents the nucleus and the latter the atom. Emphasise that the space in the atom is mostly empty and that the electrons are distributed in this empty space.

### 3.4 Atomic number

Having discussed that the mass of the atom is concentrated in the nucleus, it is now necessary to show how many protons there are in the nucleus. Point out Moseley's work on X-rays of elements which showed that atomic nuclei contain a whole number of protons, which is equal to an ordinal number known as Atomic Number, which represents the position of the element in the periodic table. Draw the attention of the students to the atomic numbers of a few elements making use of the periodic table. The atomic number is also the same as the number of electrons in a neutral atom.

### 3.5 The neutron

With the help of atomic number lead the students to guess how many protons there should be in a helium atom. If the mass number of a proton (which is nothing but a hydrogen atom which has lost the electron) is one atomic mass unit, let the students guess that the mass number of helium atom is two. Point out to them the actual mass number of helium. Make the students speculate what the extra mass is due to. It may be due to some other particles present in the nucleus which possess large mass and no charge. If they were to bear charge, the electrical neutrality of the atom would be disturbed. Point out that the existence of such particles in an atom was first predicted by Rutherford in 1920 and later discovered by Chadwick (in 1912) who called them neutrons. The mass of a neutron is practically the same as that of a proton. Since electrons have a negligibly small mass, the mass of atom is practically due to the protons and neutrons. The mass of a nucleus is generally represented by mass number ( $A$ ). It gives the total number of neutrons and protons. The atomic number ( $Z$ ), on the other hand, gives the number of protons. The number of neutrons in the nucleus, therefore, would be the difference between the mass number ( $A$ ) and atomic number ( $Z$ ).

Providing the data of mass and atomic numbers of a few atoms, the students may be asked to find out how many protons and neutrons are present in them.

### 3.6 Isotopes

Explain the term isotope with the help of examples given in the Table 3.3 of the text. Emphasise that due to the existence of isotopes the atomic masses of elements are not simply the arithmetic means of the masses of different isotopes but they are their weighted means. The idea of a weighted mean is made clear in the following problem which may be worked out in the class

#### *Problem*

Natural chlorine consists of 75.1% of chlorine-35 and 24.6% of chlorine-37 atoms. What is the atomic mass of natural chlorine? 1 mole of natural chlorine contains 0.754 mole of chlorine-35 and 0.246 mole of chlorine-37 atoms.

Weight of 0.754 mole of chlorine-35 atoms

$$\begin{aligned} &= 0.754 \text{ mole} \times \frac{35 \text{ g}}{\text{mole}} \\ &= 26.4 \text{ g} \end{aligned}$$

Weight of 0.246 mole of chlorine-37 atoms.

$$\begin{aligned} &= 0.246 \text{ mole} \times \frac{37 \text{ g}}{\text{mole}} \\ &= 9.10 \text{ g} \end{aligned}$$

$$\begin{aligned} \text{Weight of 1 mole of natural chlorine atoms} &= (26.4 + 9.1) \text{ g} \\ &= 35.5 \text{ g} \end{aligned}$$

Therefore, atomic mass of natural chlorine is 35.5

Make the students realise that the fractional atomic mass of any element is due to the existence of isotopes.

It is also necessary to bring out clearly the difference between the mass number and atomic mass. While the former refers to the number of protons plus neutrons in an atom, the latter refers to the weighted average mass of isotopes of an element as they occur in nature. For example, the mass numbers of isotopes of chlorine  $^{17}\text{Cl}^{35}$  and  $^{17}\text{Cl}^{37}$  are 35 and 37. The atomic mass of chlorine on the other hand is the weighted mean of these isotopes, taking into consideration the abundance of these isotopes in chlorine gas (see problem). The atomic mass of chlorine is 35.5.



### 3.7 Arrangement of electrons around the nucleus

So far the arrangement of protons and neutrons has been discussed. The students may be asked to guess where the electrons in an atom might be situated. Present the Bohr's model of the atom in which the electrons move in shells around the nucleus. Emphasise that only certain shells are possible. The term principal quantum number ( $n$ ) is introduced only with reference to the number of shells starting from the nucleus. It has however much greater significance which may not be discussed. The energy aspects of the shells may be avoided at present. The maximum number of electrons that can be accommodated in each shell may be given as  $2n^2$ . Emphasise that this is only an empirical rule. As there is nothing that contradicts it, it is accepted as a fact.

Make the students write the electronic configurations of a few atoms providing them with the atomic numbers.

**Table 3.1 Some common isotopes**

Element	Isotope	Atomic number	Protons	Neutrons	Mass No.	% age abundance	Atomic mass or weight
Hydrogen	Hydrogen-1	1	1	0	1	99.984	1.008
	Hydrogen-2	1	1	1	2	0.016	
	Hydrogen-3	1	1	2	3	Trace	
Lithium	Lithium-6	3	3	3	6	7.4	6.939
	Lithium-7	3	3	4	7	92.6	
Boron	Boron-10	5	5	5	10	18.83	10.811
	Boron-11	5	5	6	11	81.17	
Carbon	Carbon-12	6	6	6	12	98.89	12.011
	Carbon-13	6	6	7	13	1.11	
	Carbon-14	6	6	8	14	Trace	
Nitrogen	Nitrogen-14	7	7	7	15	99.64	14.001
	Nitrogen-15	7	7	8	14	0.36	
Oxygen	Oxygen-16	8	8	8	16	99.76	15.999
	Oxygen-17	8	8	9	17	0.04	
	Oxygen-18	8	8	10	18	0.20	
Chlorine	Chlorine-35	17	17	18	35	75.53	35.453
	Chlorine-37	17	17	20	37	24.47	
Uranium	Uranium-235	92	92	143	235	0.72	238.303
	Uranium-238	92	92	146	238	99.28	

## BACKGROUND INFORMATION

### Introduction

#### Complexity of the atom

Atomic concept of matter is one of chemistry's major contributions to man's understanding of the behaviour of matter. The first crude ideas about particle nature of matter came from Greek philosophers 2000 years ago. However, it was only in the beginning of the 19th century that a more meaningful and specific atomic concept was proposed by John Dalton. This marked the birth of the idea of an indivisible atom with a definite mass. Dalton's atomic theory could explain the laws of chemical combination. Because of such remarkable successes the validity of the theory was never doubted and at no time in the 19th century did the chemists appear concerned about the indivisibility of the atom. Even the most eminent chemists of those times could not imagine that these tiny atoms could have a complex structure. However, towards and after the turn of the 19th century considerable evidence was available to indicate that atoms are not the simple indivisible particles as Dalton thought, but are much more complex in their structure.

#### Faraday's experiments on electrolysis

In 1832-1833 Faraday's studies on electrolysis convincingly demonstrated the electrical nature of matter and the equivalence between chemical transformations and the quantity of electricity. The term 'electron' was first applied by Johnston Stoney in 1874 to the unit of electrical charge which he thought was associated with atoms.

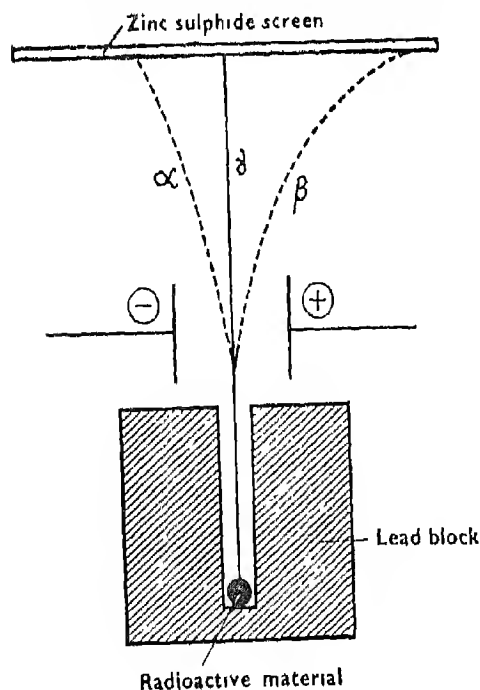
#### 3.1 The positive ions (Positive rays)

The study of electric discharge through gases at low pressures led to the discovery of cathode rays (Teachers' Guide, Chapter 2) towards the end of the 18th century. Systematic studies on cathode rays by J. J. Thomson revealed that these rays are a stream of negatively charged particles which were later given the name electrons. Their production is independent of the nature of the cathode or of the gas filling the discharge tube. They are also obtained when metal filaments are heated (thermionic emission) or when metals like sodium, potassium and cesium are exposed to light (photoelectric effect). This led to the recognition of electron being a constituent of all kinds of matter. Millikan (1909) determined the charge carried by an electron. It was shown to be  $1.6 \times 10^{-19}$  coulombs. The electron was shown to have a negligible mass, about  $1/1837$  of the mass of a hydrogen atom.

The discovery of the electron as a subatomic particle points out the possible existence of positively charged particles because the atom is electrically neutral. Goldstein discovered positive rays in a cathode ray tube experiment. It was observed that the nature of positive rays depended on the nature of the gas taken in the tube. The lightest positive particle detected had a mass which was the same as that of the hydrogen atom, the lightest atom known and a charge equal to and opposite in sign to that of the electron. This particle was called 'proton' and was considered to be the positive subatomic particle.

### 3.2 Radioactivity

Further evidence that atoms consist of electrically charged particles comes from the phenomenon of radioactivity. Towards the end of the 19th century, a French physicist, Becquerel was studying the fluorescence of uranium minerals (emission of light of various colours when exposed to sunlight, cathode rays or even X-rays). He placed a piece of uranium mineral (pitchblende,  $\text{U}_3\text{O}_8$ ) on a box containing photographic plate and left it in sunlight for a few hours. On developing the plate, he found it to be fogged, just as expected. He thought that sunlight caused the mineral to give off rays which could penetrate the cardboard box. Once, while waiting for sunshine he placed a piece of mineral on a box of photographic plate and put it aside in a desk drawer. These photographic plates were also found fogged. This was unexpected since the plate was stored in darkness. He thus made the accidental discovery that uranium minerals emit some radiation all the time; they do not need sunlight or X-rays for activation. He used the word *radioactive* for the substances which emit radiations of their own accord. The phenomenon was termed *radioactivity*. The uranium mineral was suspected to contain at least one other radioactive substance in it, for after the extraction of uranium the mineral residues showed radioactivity. Marie Curie and her husband Pierre Curie in their celebrated experiment lasting four years (1898–1902) separated a new metal polonium and yet another radioactive element radium which was two million times more active than uranium. Several other radioactive substances have been discovered since then.

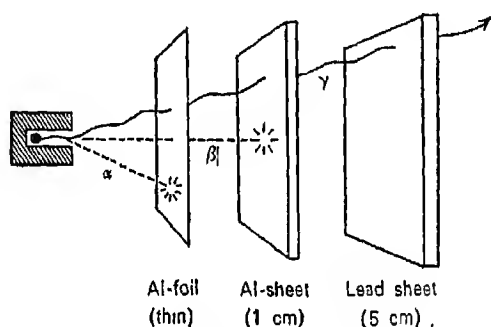


Further study of radioactivity has shown that this property is independent of the state of chemical combination of the element and that the radiation emitted by a radioactive substance consists of three kinds of rays. Rutherford showed that when a strong electric field is applied perpendicular to the direction of these radiations some were found to be deflected towards the negative plate and some towards the positive plate and some others did not suffer any deflection. The rays deflected towards the negative plate were called  $\alpha$ -rays. The magnitude of deflection was small in this case (Fig. 3.1)

Fig. 3.1 Deflections of  $\alpha$ ,  $\beta$  and  $\gamma$  rays in an electric field

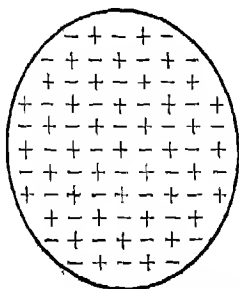
These were found to consist of positively charged particles, each particle having a mass of 4 amu. These particles were shown to be helium nuclei. The rays deflected towards the positive plate were given the name  $\beta$ -rays. The magnitude of deflection was quite large. These rays were found to consist of light negatively charged particles. These were shown to be electrons. The undeflected rays are called  $\gamma$ -rays. These are similar to light rays but are of much shorter wave length (shorter than even the X-rays). Hence they are extremely energetic and highly penetrating. The  $\alpha$ -rays because of their heavy mass cannot penetrate through matter of large thickness, whereas  $\beta$ -rays with small mass have greater penetrating ability. The relative penetrating ability of  $\alpha$ ,  $\beta$  and  $\gamma$ -rays is schematically shown in Fig. 3.2,

$\alpha$ -rays are stopped by a thin Al foil,  $\beta$  particles are stopped by an aluminium sheet of 1 cm thick.  $\gamma$ -rays can be stopped only by a sheet of lead thicker than 5 cm (Fig 3.2).



*Fig. 3.2 Penetrating power of  $\alpha$ ,  $\beta$  and  $\gamma$  rays*

Thus experiments on electrolysis revealed a relation between the quantity of electricity and the amount of material deposited. Discovery of radioactivity demonstrated the complexity of atoms. The discharge tube experiments indicated the presence of protons and electrons. Attempts were then made to understand how these different particles were arranged in the atom, in other words, the internal structure of the atom.



### 3.3 The atomic nucleus

J. J. Thomson proposed that electrons and protons in the atom are distributed uniformly as shown in Fig. 3.3.

*Fig. 3.3 Thomson's model of the atom*

In 1911 Rutherford performed a unique experiment. He allowed a narrow beam of  $\alpha$ -particles from a radioactive substance to strike a very thin foil (about 0.00004 cm thick) of gold (Fig. 3.4) and placed a screen coated with zinc sulphide at different angles around the gold foil to see how the  $\alpha$ -particles were scattered as they passed through the foil. The  $\alpha$ -particles are detected by the flash of light produced when they hit the fluorescent screen. If protons were arranged uniformly in an atom, as was proposed by Thomson, almost all the  $\alpha$ -particle should be deflected. Surprisingly Rutherford observed that a large number of the  $\alpha$ -particles went through the foil without suffering any deflection; however, some were deflected from their path by a certain small angle and about one in 20000  $\alpha$ -particles was deflected through angles of  $90^\circ$  or greater.

He made the following inferences from his observations. (1) Atom consist of large empty space since most of the  $\alpha$ -particles went through the foil undeflected. (2) The  $\alpha$ -particles which were strongly deflected must have approached a point which is highly concentrated with positive charge and got strongly repelled. Since very few particles suffered deflection through angles greater than  $90^\circ$ , the volume occupied by the positive charges must be quite small as compared to the total volume of the atom.

On the basis of these inferences, Rutherford proposed the *Nuclear model* of the atom. He suggested that the atom is mostly empty space in which electrons are located and that all the positive charges as well as most of the mass of the atom are concentrated in a very tiny portion within the atom which he called the *Nucleus*. It was estimated that the diameter of the nucleus is of the order of 0.000000000001 cm or  $1.0 \times 10^{-13}$  cm.

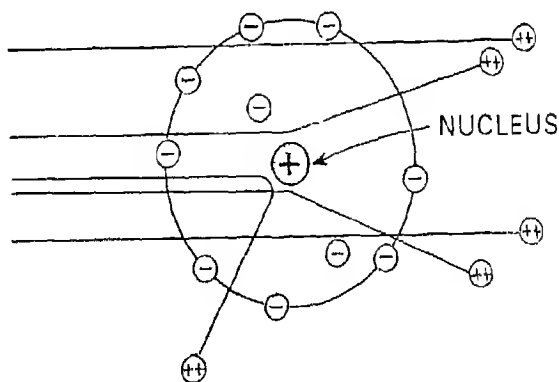


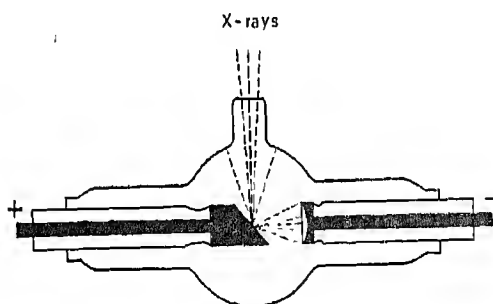
Fig. 3.4 Scattering of  $\alpha$ -particles

The diameter of single atom is determined to be of the order of 0.00000001 cm or  $1 \times 10^{-8}$  cm.

From this it is evident that the nucleus is  $10^5$  times smaller than the atom. This minuteness is difficult to imagine. An analogy might help. If the atom had a diameter of 1 Km, its nucleus would have a diameter of the order of only 1 cm. Or if the atom were the size of an average house, the nucleus would correspond to just a minute pinhead in its centre. In spite of its very small volume it contains most of the mass of the atom. Electrons which surround the nucleus furnish very little of the atomic mass.

### 3.4 Atomic number

In 1913 Moseley studied the X-ray spectra of elements from aluminium to gold in the periodic table. X-rays are produced when a beam of fast moving electrons (cathode rays) strike a metal. A discharge tube was used for this purpose making the metal the anticathode and the X-ray spectrum was obtained. A typical X-ray tube is shown in Fig. 3.5. Moseley observed a regularity in the wave length of the emitted



*Fig. 3.5 X-ray tube*

X-rays. There was a constant increment in the square root of the frequency ( $\nu$ ) of the individual X-radiation emitted as we pass from one element to the other in the increasing order of atomic number. The idea of atomic number had been suggested much earlier to indicate the ordinal number of the element in the periodic table. Moseley suggested that the square root of the frequency of characteristic X-radiation of an element is directly proportional to the atomic number of the element.  $\nu = (Z-b)^2$ . number ( $Z$ ) of an element is equal to the number of protons in the nucleus of the atom of the element. It is also equal to the number of electrons outside the nucleus of a neutral atom. Since the physical and chemical properties of an element are determined by the number of electrons in its atom, the atomic number is a characteristic property of the element.

### 3.7 Arrangement of electrons around the nucleus

According to Rutherford's nuclear model, the atom consists of protons and neutrons packed into a very small, compact nucleus which is surrounded by electrons

occupying a relatively much larger space. Since electrons in an atom determine its chemical behaviour, the arrangement of these electrons is of primary importance. Niels Bohr and others postulated that the electrons were in rapid motion in circular orbits around the nucleus. Because of their rapid motion, the electrons of the atoms possess some energy, the kinetic energy. They also possess potential energy due to attraction by the nucleus and the repulsion of other electrons. These two types of energies together determine the energy of any electron in an atom. Do all electrons in an atom have the same energy? In other words, do all the electrons in an atom occupy the same region around the nucleus?

Bohr proposed that the electron of hydrogen can exist only in certain orbits or shells around the nucleus. The energy of an electron is assumed to be unaltered as long as it stays in a particular orbit or shell. He called such shells the stationary shells or energy shells or energy levels. The stationary shells or energy levels are numbered 1,2,3 etc. starting from the nucleus. The numbers are called the principal quantum numbers ( $n$ ). The shells are also designated as K,L,M, etc shells. With increasing distance from the nucleus the radius of the shell increases and the energy of electrons in shell increases. It was shown that the energy of an electron in a given shell of hydrogen that atom can be represented by

$$E_n = - \frac{313,6}{n^2} \text{ kcal/mole}$$

where  $E_n$  is the energy of the electron in the  $n^{\text{th}}$  shell and  $n$  the shell number.

The maximum number of electrons that can occupy any shell is given by  $2n^2$ , where  $n$  is the principal quantum number of the shell in question (Table 3.4 of Text). The K shell (principal quantum no. 1) can accommodate two electrons, the L shell (principal quantum no. 2) can accommodate 8 whereas the M and N shells accommodate 18 and 32 electrons respectively.

<i>Shell</i>	<i>Principal quantum No.</i>	<i>No. of electrons</i>
K	1	2
L	2	8
M	3	18
N	4	32

The electrons, as far as possible, tend to go to the lowest energy shell available. When this is filled completely the electrons go to the next higher energy level till that is completely filled. Of the six electrons of carbon ( $Z = 6$ ) two would occupy the K shell and four would go to the L shell.

	K	L
C	2	4

This rule works for elements up to argon ( $Z = 18$ ). However, it fails in the case of heavier elements and hence it should be modified. For example, potassium ( $Z = 19$ ) which should have the electron configuration

	K	L	M	N
K :	2	8	9	-

has actually the configuration

K :	2	8	8	1
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How can this be explained ?

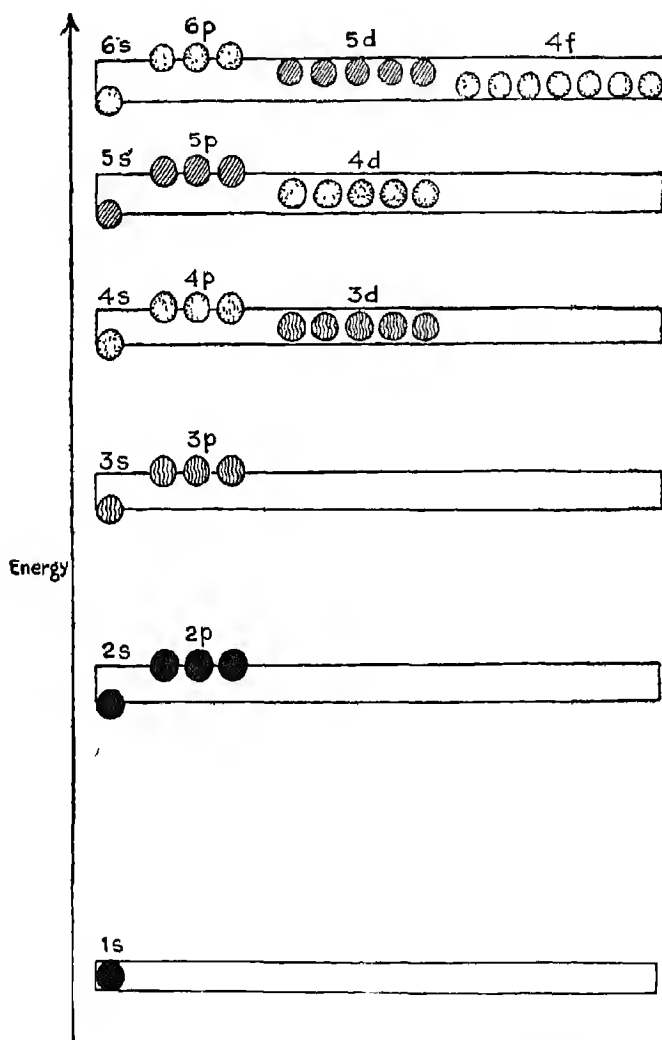


Fig. 3.6 The energy level sequence of atoms (except hydrogen atom)



Spectroscopic investigations indicated that each shell (or level) contains subshells (or sublevels). The number of sublevels in a main energy level is equal to its principal quantum number,  $n$ . The level ( $n = 1$ ) has one sublevel, called  $s$ ; level 2 has two sublevels called  $2s$  and  $2p$ , and level 3 has three sublevels;  $3s$ ,  $3p$  and  $3d$ . Sublevels belonging to the same main level have the same energy in the case of the hydrogen atom. But this is not so in the case of atoms of other elements. For example, in main level 2, the  $2p$  sublevel has a slightly higher energy than  $2s$ . In the main level 3, the  $3d$  level has a higher energy than the  $3p$  level, which in turn has higher energy than the  $3s$  level. The energy sequence of the sublevels of main level 3 is thus  $3d > 3p > 3s$ . A schematic energy level sequence is shown in Fig 3.6. It is of interest to note that the  $3d$  level is higher in energy than the  $4s$  level; the electrons would enter the  $4s$  level before entering the  $3d$  level. After the completion of the  $4s$  level the  $3d$  level starts filling.

## Quiz

Each question is followed by four answers. One of them is correct. Indicate the correct answer in the bracket provided.

1. The mass of positive particles formed in a discharge tube ( )  
(a) depends on the nature of the gas  
(b) does not depend on the nature of the gas  
(c) depends on the nature of the electrode used  
(d) depends on the potential used.
2. A positively charged particle having unit atomic mass is formed in a discharge when the gas used is ( )  
(a) helium (b) oxygen  
(c) nitrogen (d) hydrogen.
3. A substance capable of emitting invisible rays which affect a photographic plate (even when wrapped in a black paper) is ( )  
(a) sodium salt (b) potassium salt  
(c) thorium salt (d) magnesium salt.
4. An example of a radioactive element is ( )  
(a) aluminium (b) calcium  
(c) iodine (d) polonium.
5. From Rutherford's experiment on the scattering of  $\alpha$ -rays we conclude that ( )  
(a) the mass and charge is uniformly distributed throughout the atom  
(b) the mass alone is concentrated in a small region in the atom  
(c) the mass and positive charge are concentrated in a small region in the atom  
(d) positive charge alone is concentrated in a small region in the atom.

6. Radioactive elements disintegrate ( )
- (a) when exposed to radiation
  - (b) when heated
  - (c) spontaneously
  - (d) when subjected to a magnetic field.
7.  $\alpha$ -rays are
- (a) positively charged      (b) negatively charged
  - (c) neutral      (d) singly charged helium ions.
8. The atomic number of an element is equal to the ( )
- (a) protons in the nucleus
  - (b) neutrons in the nucleus
  - (c) sum of protons and neutrons in the nucleus
  - (d) difference between the protons and neutrons in the nucleus.
9. The atomic number of sodium is 11 and its mass number is 23. Hence the number of neutrons in its nucleus is ( )
- (a) 12      (b) 11
  - (c) 23      (d) 13.
10. Oxygen has an atomic number 8 and a mass number 16. The number of neutrons present in the nucleus of an oxygen atom is ( )
- (a) 4      (b) 2
  - (c) 6      (d) 8.
11. The mass number of an element is ( )
- (a) the number of protons present in it
  - (b) the number of neutrons present in it
  - (c) the sum of the number of protons and neutrons present in it
  - (d) equal to the atomic number.

12. The atomic nucleus consists of ( )  
(a) protons and electrons (b) protons only  
(c) protons and neutrons (d) electrons and neutrons.
13. A particle having two positive charges and a mass number 4 is called ( )  
(a) proton (b) helium atom  
(c)  $\alpha$ -particle (d)  $\beta$ -particle.
14.  $\beta$ -rays are a stream of ( )  
(a) helium nuclei  
(b) electrons  
(c) neutrons  
(d) protons.
15. The particle obtained when a hydrogen atom loses an electron is ( )  
(a) neutron (b) proton  
(c)  $\alpha$ -particle (d)  $\beta$ -particle.
16. The atom is larger than the nucleus by  
(a) 100 times (b) 1,000 times  
(c) 10,000 times (d) 100,000 times.
17. Isotopes of an element have ( )  
(a) different atomic numbers but the same mass number  
(b) different mass numbers and different atomic numbers  
(c) different mass numbers but the same atomic number  
(d) same mass number and same atomic number.
18. The maximum number of electrons in the L shell of an atom is  
(a) 2 (b) 8  
(c) 16 (d) 32.

19. The maximum electron population of the M shell of an atom is ( )
- (a) 3 (b) 6
- (c) 9 (c) 18.
20. The atomic number of chlorine is 17 and the mass number of one of its isotopes is 35. The nucleus of this isotope consists of ( )
- (a) 17 protons and 18 neutrons
- (b) 18 protons and 17 neutrons
- (c) 35 protons and no neutrons
- (d) no protons and 35 neutrons.

# How Atoms Combine

## INSTRUCTION MATERIAL

### Introduction

As we know, there are about 104 elements which constitute all matter. It is known that the atoms of elements combine in different characteristic ways to form a variety of chemical compounds, each having specific properties. In this chapter we shall study the different ways in which atoms combine to form molecules.

### Outline

1. Combination between atoms ( 4.1 )
2. Ionic bond ( 4.2 )
3. Covalent bond ( 4.3 )
4. Polarity of a covalent bond ( 4.4 )

### Concepts

1. Atoms of elements during combination tend to acquire the stable electronic configuration of the noble gas atoms – either by losing or gaining or sharing some of their electrons.
2. Transfer of electrons from one atom to another leads to the formation of an ionic bond.

3. Sharing of electrons between atoms results in the formation of a covalent bond.
4. Unequal sharing of electrons by the atoms results in polar character of the covalent bond.

### SCHEDULE

Lesson	Experiment	Postlab	Demonstration	Text covered
1	—	—	—	4.1
2	—	—	—	4.2
3	—	—	—	4.3
4	—	—	—	4.4
5	—	—	—	Review and quiz

### Development

#### 4.1 Combination between atoms

Recapitulate the important ideas of the previous chapter, particularly, the distribution of electrons in different shells. Write down the electronic configurations of the atoms of noble gases, alkali metals and halogens. Recall that atoms of noble gases are known to be remarkably stable and show little tendency to combine with other elements. Atoms of alkali metals and halogens, on the other hand, are extremely reactive. They have incomplete outermost shells. Mention that atoms of alkali metals and halogens enter into chemical reactions in an attempt to acquire stable configuration of the noble gas type.

#### 4.2 Ionic bond

Ask the students to write down the electronic configurations of sodium and chlorine atoms. Tell them that the electrons in the outermost shell are the ones which take part in chemical reactions and are called the valence electrons. For simplicity in writing the chemical reactions it is useful to represent the nucleus and inner electrons of the atom by the symbol of the atom and to show only the outermost electrons by dots. Elicit from the students as to how the atom in question can acquire stable configuration of a noble gas. Discuss the formation of  $\text{NaCl}$ ,  $\text{MgO}$ ,  $\text{MgCl}_2$  and  $\text{MgS}$ . Emphasise that ions formed are held together by electrostatic forces that are quite strong.

Mention some of the properties of ionic compounds at this stage. Ionic compounds are fairly hard and brittle. They are soluble in water and insoluble in organic solvents. Their melting and boiling points are high. They conduct electric current when they are in a fused state or in aqueous solution.

### **4.3 Covalent bond**

Mention that another way of attaining the stable configuration of the noble gas type is by mutual sharing of electrons. Discuss the formation of molecules of hydrogen, chlorine, hydrogen chloride and water. Mention some of the properties of covalent compounds at this stage.

Covalent compounds are relatively soft. They have low melting and boiling points. They are not very soluble in water, but are usually soluble in organic solvents like carbon tetrachloride, alcohol etc. They are poor conductors of electricity.

### **4.4 Polarity of a covalent bond**

Discuss how unequal sharing of electrons leads to polarity of a covalent bond with examples of hydrogen chloride and water. Emphasise that unequal sharing of electrons arises due to the difference in the electron attracting ability of the atoms. Introduce the idea of polar and non polar molecules with suitable examples.



## BACKGROUND INFORMATION

### Why do atoms combine ?

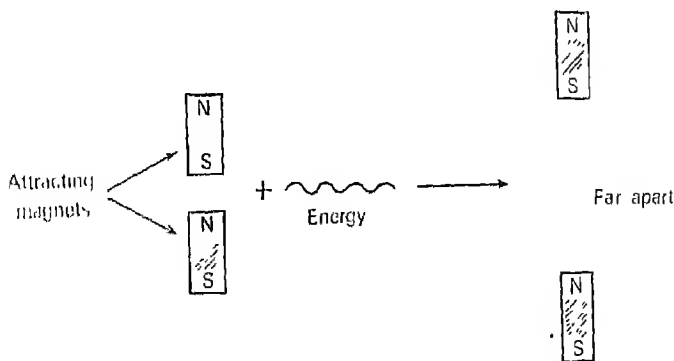
The development of modern theory of atomic structure has considerably revolutionised our concepts about chemical bonding. Before discussing the modern concepts of chemical bonding let us first understand some guiding principles of general applicability which will be found useful in deciding the stability of a chemical bond.

#### (1) Systems tend to decrease their energy

It is well known that objects tend to attain a state of minimum energy. For instance, a red hot iron bar will radiate its energy to the surroundings and will thus cool down. An object let loose from a height will fall to the earth, thus losing its potential energy. It is for the same reason that water flows down hill. That in all these cases energy has been lost is evident from the fact that to reverse these changes work has to be done i.e. energy has to be supplied. Thus, we realise an important principle that all systems tend to decrease their energy as and when there is a chance for them to do so.

#### (2) Attractive interaction leads to decrease of potential energy while repulsive interaction leads to an increase

We know that the opposite poles of two magnets can approach each other due to the attractive interaction between them (Fig. 4.1). No work need be done in bringing the opposite poles of a magnet near each other. However, energy is required to pull these poles apart. Thus we find that when two objects attract each other work has to be done to separate them. This means that as a result of attraction, the object has lost some energy which had to be supplied to separate them. We, therefore realise that attractive interaction leads to a decrease of energy of the system.



*Fig. 4.1 Energy spent in separating the poles of two magnets*

On the other hand, if we wish to bring north poles of two magnets near each other, work has to be done to overcome the force of repulsion between them. This means that the system in which the two north poles are brought together has more energy. In other words, repulsive interaction leads to an increase of energy.

We thus formulate an important principle that attractive interaction leads to a decrease of energy and repulsive interaction leads to an increase of energy of the system.

### The electron-nucleus system

On the basis of the above analogy it can be inferred that when an electron approaches the nucleus (thereby increasing the attractive interaction), energy is released. Hence the total energy of the atom decreases (Fig. 4.2). Conversely to move the electron away from the nucleus, energy has to be supplied. Hence the total energy of the atom increases.

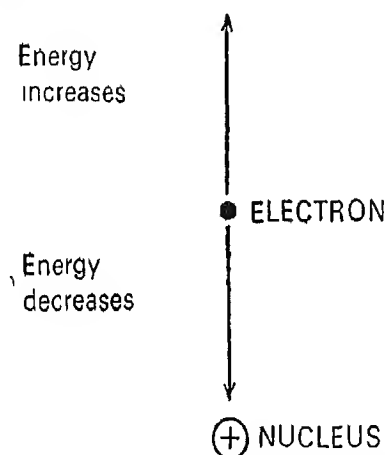


Fig. 4.2 Energy changes during the interaction between electron and nucleus

### Stability of a chemical bond

Let us consider the case of a molecule of hydrogen formed by the combination of two atoms of hydrogen. It is known that energy is released when two atoms combine to form a molecule (Fig. 4.3). In other words a molecule of hydrogen has lower

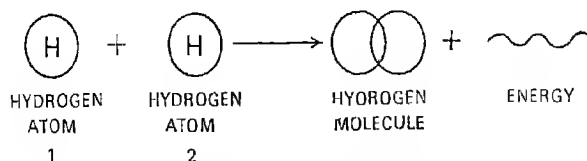


Fig. 4.3 Formation of a molecule of hydrogen

energy than the separate atoms. What causes this energy lowering? Energy changes take place when charged particles in the atom interact. Let us see what happens when two hydrogen atoms approach each other. Each hydrogen atom has a nucleus which is positively charged and an electron moving around it. In an isolated atom the electron is under the influence of its own nucleus. But as the two atoms (atoms 1 and 2) come closer, the electron of atom 1 will be attracted to the nucleus of atom 2 and vice versa. Both electrons spend a great deal of their time in the region between the two nuclei. In this region each electron is attracted to both the nuclei. The atoms are therefore drawn closer and energy is released. Consequently the potential energy is decreased. Such an arrangement is energetically more stable than two separated atoms. At a particular distance between the two nuclei the attractive forces are greater than the repulsive process and so a stable chemical bond is formed. If the nuclei try to approach closer than this distance, the repulsive forces increase and the system becomes unstable.

## 4.2 Ionic bond

### Energy and ionic bonding

Ionic bonds are formed by the transfer of electrons from the outer energy level of an electropositive atom to the outer energy level of an electronegative atom.



Let us examine the energy changes involved in the formation of crystalline NaCl. This reaction is exothermic. But we are interested in the step-wise energy changes which lead to the final evolution of heat.

Firstly, solid sodium must be dissociated into free atoms. This requires the absorption of energy, the heat of sublimation of sodium. Secondly, molecular chlorine must be dissociated into free atoms. This requires the absorption of energy, i.e. heat of dissociation of chlorine.

Thirdly, the outer 3s electron (valence electron) of sodium must be removed from the atom. The energy needed to do this, called the ionization energy, is absorbed.

Fourthly, as the valence electron (3s) of the sodium atom is transferred to the outer energy level (3p) of the chlorine atom, energy is liberated in this process. This energy is the electron affinity of chlorine.

Fifthly, after the sodium and chloride ions have been formed by the transfer of the electron, electrostatic attraction will draw them together to form an ion pair. In a large sample, the ion pairs will become arranged into a crystal lattice, liberating energy called the crystal, or lattice energy.

The reaction between sodium and chlorine occurs only because the energy released is greater than the energy absorbed. The product is therefore in a lower energy state than the reactants, and thus relatively more stable. This is true for the formation of all ionic bonds.

These energy changes per mole are summarised in Fig. 4.4.  $E_1$  = the heat of sublimation of sodium;  $E_2$  = the heat of dissociation of chlorine;  $E_3$  = the ionization energy of sodium;  $E_4$  = the electron affinity of chlorine; and  $E_5$  = the lattice energy of NaCl.

$$\begin{aligned} E_1 &= +26 \text{ KCal.} & \frac{1}{2} E_2^* &= +29 \text{ KCal,} \\ E_3 &= +118 \text{ KCal} & E_4 &= -87 \text{ KCal,} \\ E_5 &= -184 \text{ KCal} \end{aligned}$$

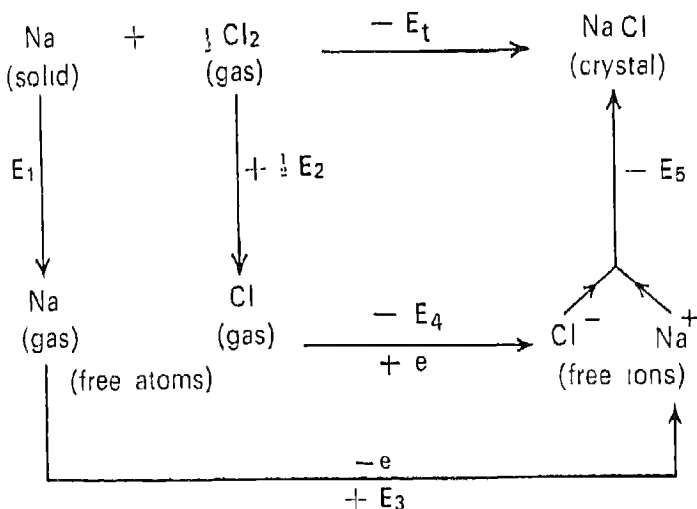


Fig. 4.4 Formation of sodium chloride crystal lattice

Notice that plus (+) sign is used to indicate energy absorbed and a minus (—) sign for energy liberated.  $E_t$  is the total energy liberated.

$$E_t = 26 + 118 + 29 - 87 - 184 \text{ Kcal/mole} = -98 \text{ Kcal/mole.}$$

This calculated value is found to be the same as the experimentally determined value.

\*Heat of dissociation defined per mole. Therefore  $\frac{1}{2} E_2$  is the energy involved in the cycle.

As the ions in an ionic solid are rigidly held by strong forces of attraction, ionic solids are usually hard and brittle.

Ionic compounds are soluble in polar solvents like water and insoluble in non polar solvents like carbon tetrachloride.

Let us now consider how an ionic compound dissolves in a polar solvent like water. When an ionic solid is placed in water, the ions of the solid attract the polar solvent molecules. The negative ends of the water molecule are bound to the positive ion and the positive ends to the negative ions. This weakens the inter-ionic forces and eventually the ions come off the crystal lattice. The separated ions are surrounded by a number of molecules of water.

In ionic solids, the ions of opposite charge are held by strong forces of electrostatic attractions. A large amount of energy is needed to pull apart the oppositely charged ions. Hence their melting and boiling points are high.

The ionic compounds if melted or dissolved in water conduct electric current. This is because the ions responsible for conduction are free to move only when melted or dissolved in water. In the solid state, on the other hand, the ions are held rigidly in the lattice and are not free to move.

### **4.3 Covalent bond**

As the molecules in covalent compounds are held by weak forces, they can be easily pushed around from one place to another. Hence they are relatively soft.

Unlike in the ionic solids, the total intermolecular attractions in covalent compounds are small. So less energy is needed for pulling the molecules apart. The melting and boiling points of covalent compounds are therefore relatively low.

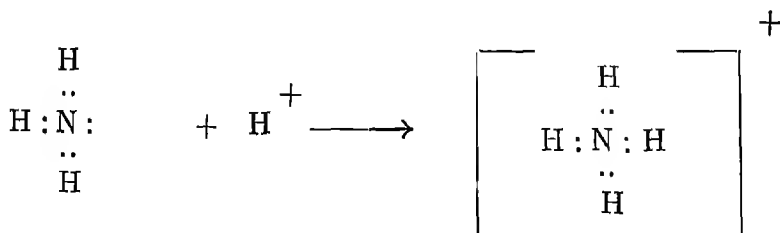
Covalent compounds are usually poor conductors of electricity since the basic particle in covalent compounds is not an ion but a molecule e. g., carbon disulphide, benzene etc. The covalent compounds do not dissolve in polar solvents like water but dissolve in non polar solvents. For example, NaCl dissolves in water but not in benzene. A covalent compound like naphthalene dissolves in benzene but not in water.

**Table 4.1 Melting and boiling points of some ionic and covalent compounds**

Substance	Melting point (°C)	Boiling point (°C)
Sodium chloride	804	1413
Potassium iodide	723	1420
Potassium bromide	730	1380
Hydrogen chloride	-112	- 83.7
Oxygen	-218.8	-183.0
Nitrogen	-209.9	-195.8
Chlorine	-101	- 34.1
Iodine	113.5	184.4
Carbon tetrachloride	9	189

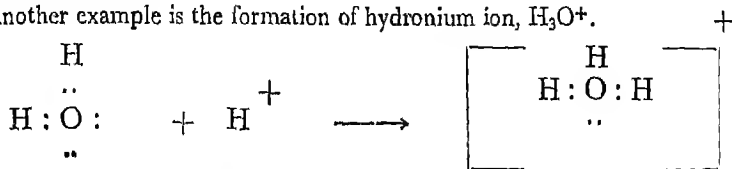
### Coordinate covalent bond or dative linkage

We know that a covalent bond consists of a pair of electrons shared by two atoms. One of the electrons in the pair is supplied by one bonding atom and the other electron is supplied by the other atom. In some covalent bonds, however, one of the atoms supplies both the electrons. For example, consider the formation of ammonium ion,  $\text{NH}_4^+$ . A molecule of ammonia,  $\text{NH}_3$ , reacts with a hydrogen ion,  $\text{H}^+$ , to form an ammonium ion. This change is represented as



In this reaction, nitrogen atom supplies both the electrons (nitrogen has an extra lone pair of electrons not involved in the bond formation in ammonia) and is known as the donor atom. The hydrogen ion,  $H^+$ , (called acceptor) readily shares the lone pair of electrons offered by the nitrogen atom and in so doing, achieves a helium like stable configuration. The chemical change in this reaction consist of the sharing of the lone pair of electrons of the nitrogen atom with the hydrogen ion. The product of this change is the ammonium ion which has one positive charge. When once the ammonium ion is formed, the four  $N - H$  bonds of the ion become identical and indistinguishable. A covalent bond formed in this way is called a coordinate covalent bond. The compounds containing such bonds are known as coordination compounds.

Another example is the formation of hydronium ion,  $H_3O^+$ .



### How many bonds can an atom form?

Electron transfer or sharing may continue till a maximum of eight electrons or four pairs, surround an atom. In hydrogen sharing is, of course, limited to one pair. The application of this is conveniently illustrated with a few examples using electron dot formulas.

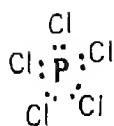
A molecule of chlorine has two atoms of chlorine. Each atom of chlorine has seven electrons in its outer most shell and the formation of a molecule of chlorine can be represented as shown below indicating that one covalent bond is formed.



It is seen that each chlorine atom is surrounded by four pairs of electrons, one pair is the bond pair and the other three are the unshared pairs. Studies have shown that further bond formation in the chlorine molecule does not take place because an 8-electron arrangement is a stable arrangement.

In water,  $H_2O$ , oxygen has 6 electrons and the two hydrogen atoms contribute one electron, each making a total of eight electrons in the outermost shell of oxygen, thus completing an octet.

However, the octet rule strictly applies to the elements of the second short period of the periodic table where, a maximum of four bonds can be formed through sharing. There are, however, a number of molecules where 5 or 6 covalent bonds are formed by the central atom. Such a behaviour is exhibited by elements such as phosphorus and sulphur.



Phosphorus (V) chloride



Sulphur (VI) fluoride

Fig. 4.6 Molecules of  $\text{PCl}_5$  and  $\text{SF}_6$

#### 4.4 Polarity of a covalent bond

##### *Electronegativity*

In order to understand the nature of a chemical bond there must be a way of comparing the electron attracting powers of the atoms concerned in the chemical bond. This can be done in terms of electronegativity of the element. The tendency of an atom to attract electrons to itself when combined in a compound is termed the electronegativity of the atom. The relative electronegativities of some elements are given in Table 4.2.

**Table 4.2 Relative electronegativities (Pauling's scale)**

Element	Relative electronegativity	Element	Relative electronegativity
F	4.0	Si	1.8
O	3.5	Sn	1.7
N	3.0	Al	1.5
Cl	3.0	Mg	1.5
Br	2.8	Li	1.0
C	2.5	Ca	1.0
S	2.5	Na	0.9
I	2.4	K	0.8
H	2.1	Rb	0.8
P	2.1	Cs	0.7

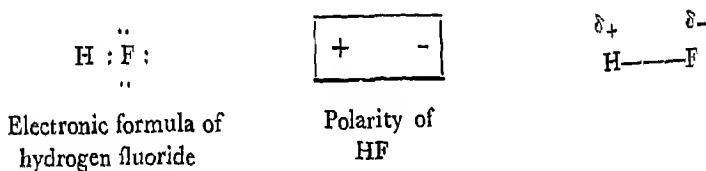


Non metals have high electronegativities compared to metals. Fluorine is the most electronegative element and cesium the least electronegative.

The greater the difference between the electronegativities of the bonding atoms, stronger is the bond formed and stabler is the resulting compound. Thus the bond in HF is very much stronger than the bond in HI. Consequently hydrogen fluoride is a stabler compound than hydrogen iodide. In the same way ammonia,  $\text{NH}_3$ , is stabler than phosphine,  $\text{PH}_3$ .

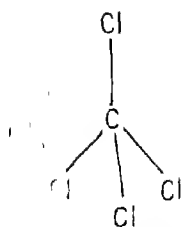
### Polar and nonpolar molecules

When a covalent bond is formed, the two bonding atoms may share the pair equally or one of the atoms may exert a greater attraction for the electrons. Let us consider the hydrogen molecule,  $\text{H} : \text{H}$ . Here the electron pair is shared equally because the atoms are exactly alike. In the case of a molecule like HF, the sharing of the electrons is not equal. Fluorine exerts a greater pull on electrons (electronegativity of F is higher than that of H, the difference being 1.9 (4.0 - 2.1). Consequently the fluorine end of the molecule becomes negatively charged and the hydrogen end positively charged.

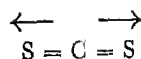


The hydrogen - oxygen bonds in water are partially ionic (electronegativity difference  $3.5 - 2.1 = 1.4$ ). Thus these are polar covalent bonds, with the oxygen being somewhat negative and the hydrogen being somewhat positive. Since the water molecule contains two such bonds unsymmetrically arranged, the water molecule as a whole shows regions of positive and negative charge and is thus polar. Likewise the nitrogen-hydrogen bonds in ammonia are polar. The molecules HF,  $\text{H}_2\text{O}$ ,  $\text{NH}_3$ , thus behave as *electric dipoles*. The product of the charge at one end and the distance between the charges is called *dipole moment*. Molecules having a dipole moment are called polar molecules. The dipole moment has a direction as well as a magnitude. It is a vector quantity.

Molecules like carbon tetrachloride,  $\text{CCl}_4$ , and carbon disulphide,  $\text{CS}_2$ , are not polar molecules although the C-Cl and C = S bonds in these molecules are polar. They have a symmetrical structure and the dipole of one bond cancels the dipole of the other. There is thus no resultant dipole moment.



Carbon tetrachloride



Carbon disulphide  
(the direction of the  
dipole is shown by  
arrow head)

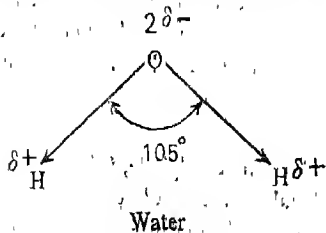
*Fig. 4.7 Nonpolar molecules*

Molecules like  $\text{CCl}_4$ ,  $\text{CS}_2$  having no dipole moment are called nonpolar molecules.

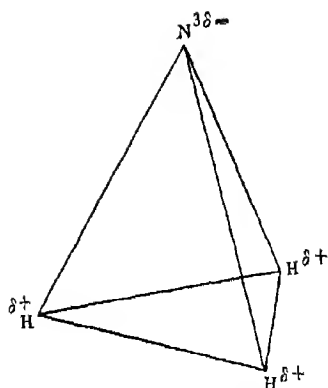
Diatomic molecules like  $\text{H}_2$ ,  $\text{O}_2$  and  $\text{Cl}_2$  are symmetrical because the two atoms are identical. So there is electrical charge symmetry and hence no dipole is observed. They are linear i.e. the two nuclei lie on a straight line. However, in triatomic molecules like  $\text{H}_2\text{O}$ ,  $\text{H}_2\text{S}$ ,  $\text{CO}_2$  and  $\text{SO}_2$  two arrangements are possible: all three atoms may lie on a straight line, that is, the molecule may be linear, or (ii) the three atoms may not be collinear, the molecule may be bent. Experiments show that carbon dioxide is a nonpolar molecule. This suggests a linear arrangement. If the arrangement were angular, the polarities of the two carbon-oxygen bonds would not cancel each other and the molecule would have been polar.



The water molecule is also triatomic. However, unlike the carbon dioxide molecule it is found to be polar. Therefore, the arrangement of the atoms in the water molecule must be angular (Fig. 4.8) and not linear. Experimentally the angle between the two oxygen-hydrogen bonds is found to be about  $105^\circ$ . Similarly the hydrogen sulphide molecule is also polar and therefore angular. The bond angle HSH is  $92^\circ$ .

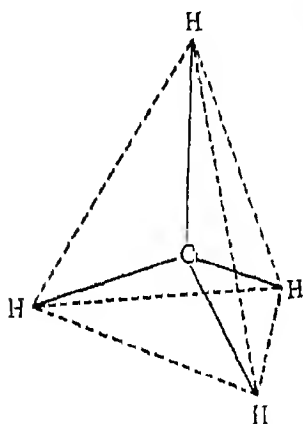


*Fig. 4.8 Structure of the water molecule.*



*Fig. 4.9 Shape of the ammonia molecule*

Consider the ammonia molecule. If the molecule were planar (i.e., all the four atoms were in the same plane) the angle between the nitrogen-hydrogen bonds would be  $120^\circ$ , and the molecule would have been nonpolar. Actually, ammonia is polar and the angle between the bonds is found to be  $107.3^\circ$ . This is possible only if the molecule is pyramidal with the nitrogen atom at the apex and the three hydrogen atoms in the base (Fig. 4.9)



*Fig. 4.10 Shape of the methane molecule*

On the other hand, consider molecules like methane,  $\text{CH}_4$ , and carbon tetrachloride,  $\text{CCl}_4$ . The carbon atom is at the centre of a regular tetrahedron and each of the C-H or C-Cl bonds is directed towards the four corners of a regular tetrahedron. It is true that each of the C-H or C-Cl bonds has a dipole. But since the four bonds in the molecule are identical and directed to four corners of a regular tetrahedron, they cancel one another resulting in a nonpolar molecule (Fig. 4.10). Thus we find that molecules having bond dipoles need not necessarily be polar in nature. The molecular shape, symmetrical or otherwise, also decides the polar nature of the molecule.

## Quiz

Each question given below is followed by four answers. One of them is correct. Indicate the correct answer in the brackets provided.

1. The element which has two valence electrons is ( )  
(a) hydrogen (b) neon  
(c) helium (d) sodium.
2. An example of the element which has one valence electron is ( )  
(a) helium (b) sodium  
(c) chlorine (d) magnesium.
3. The number of electrons in the outermost shell of an atom which indicates marked stability in elements is ( )  
(a) 3 (b) 4  
(c) 6 (d) 8.
4. The electronic configuration of aluminium (atomic number 13) is ( )  
(a) 8, 3; 2 (b) 2, 3, 8  
(c) 3, 8, 2 (d) 2, 8, 3.
5. The electronic configuration of an element with atomic number 9 is ( )  
(a) 1, 3, 5 (b) 2, 6, 1  
(c) 2, 7, 0 (d) 1, 7, 1.
6. The atom having the electronic configuration 2, 8, 6 is that of ( )  
(a) helium (b) carbon  
(c) oxygen (d) sulphur.
7. In the formation of magnesium chloride, a magnesium atom acquires the stable configuration of a noble gas by ( )  
(a) losing 2 electrons (b) losing 4 electrons  
(c) gaining 1 electron (d) gaining 2 electrons.

8. An example of an ionic compound is ( )  
(a) methane (b) hydrogen chloride  
(c) water (d) magnesium chloride.
9. An example of a covalent substance is ( )  
(a) sodium sulphate (b) magnesium bromide  
(c) hydrogen chloride (d) sodium chloride.
10. An example of a polar molecule is ( )  
(a) hydrogen (b) oxygen  
(c) ammonia (d) chlorine.
11. An example of a nonpolar molecule is ( )  
(a) water (b) carbon dioxide  
(c) ammonia (d) hydrogen chloride.
12. The atomic number of sodium is 11.  
The electronic configuration of  $\text{Na}^+$  ion is ( )  
(a) 2, 8, 1 (b) 2, 8, 0  
(c) 2, 8, 8 (d) 2, 8, 2.
13. A negative ion is formed when an atom ( )  
(a) loses an electron (b) loses a proton  
(c) gains an electron (d) gains a neutron.

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# Chemical Equations

## INSTRUCTION MATERIAL

### Introduction

A chemical equation describes a reaction indicating the reactants and products formed in terms of their symbols and/or formulae. Symbols, formulae, and chemical equations have a precise quantitative meaning. To make the equation more informative energy changes and other conditions of the reaction may also be indicated.

### Outline

1. Meaning of an equation (5.1)
2. Problems based on equations (5.2)
3. Equations showing the heat of reaction (5.3)
4. Ionic equations (5.4)

### Concepts

1. In a chemical reaction the total number of atoms and therefore the total mass is conserved.
2. In ionic reactions the charges are conserved besides mass.
3. The reactions involving heat changes are represented by thermochemical equations.

## SCHEDULE

Lesson	Experiment	Postlab	Demonstration	Text covered
1	—	—	—	5.1
2	—	—	—	5.2
3	—	—	—	5.3
4	10	Discussion	—	5.4
5	—	—	—	Review & quiz

### Development

#### 5.1 Meaning of an equation

The first equation chosen for discussion is the reaction between hydrogen and chlorine. Unfold to the students the many facts summarised in the equation one by one. Impress upon them that the equation does not merely give a list of reactants and products of the reaction taking place but it also specifies the quantities of all the species involved. The steps indicated in the text have to be developed gradually and only after making sure that the students have grasped the full meaning and significance of each term. The second equation to be discussed is that of the burning of magnesium.

Careful consideration of the detailed quantitative study made on the above chemical equations leads to relation between weights and volumes of the species involved. The branch of chemistry dealing with exact mass and/or volume relations among substances involved in a chemical reaction is called stoichiometry.

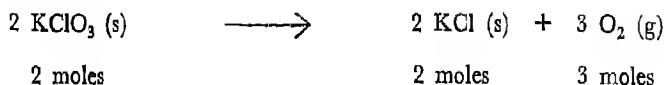
#### 5.2 Problems based on equations

In all calculations emphasis should be laid on the mole concept and it should be used often so that the student becomes more and more familiar with it.

#### Weight-weight calculations

What weight of oxygen gas will be produced from the thermal decomposition of 1.226 g of potassium chlorate? (Formula weight of  $\text{KClO}_3 = 122.6 \text{ g}$ ).

To deal with any problem in stoichiometry, we begin with the balanced chemical equation for the reaction.



First determine the number of moles of  $\text{KClO}_3$  present in 1.226 g of this compound.

$$\begin{aligned}
 \text{No. of moles of KClO}_3 &= \frac{\text{Weight of potassium chlorate}}{\text{Formula weight of potassium chlorate}} \\
 \text{in 1.226 g of this} &= \frac{1.226 \text{ g}}{1.236 \text{ g/mole}} \\
 \text{compound} &= 0.010 \text{ mole}
 \end{aligned}$$

According to the above equation, 2 moles of potassium chlorate yield 3 moles of oxygen.

$\therefore$  0.010 mole of potassium chlorate will yield

$$\frac{3 \times 0.010}{2} = 0.015 \text{ mole of oxygen}$$

But we know that 1 mole of oxygen represents 32 g of oxygen.

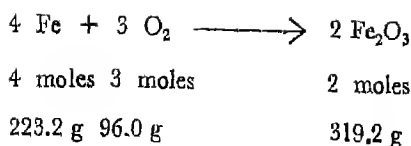
$$\begin{aligned}
 \therefore \text{Weight of 0.015 mole of oxygen} &= 0.015 \times 32 \\
 &= 0.48 \text{ g}
 \end{aligned}$$

Hence, the weight of oxygen produced from the thermal decomposition of 1.226 g of potassium chlorate is 0.48 g.

### Example 2

- How many grams of oxygen are required to oxidize 140 g of iron to iron (III) oxide?
- How many moles of iron (III) oxide are produced?  
(Atomic wt. of iron = 55.8, atomic wt. of oxygen = 16).





(a) First determine the number of moles of iron present in 140 g of iron.

$$\begin{aligned}
 \text{No. of moles of iron in 140 g} &= \frac{\text{Weight of iron}}{\text{Atomic weight of iron}} \\
 &= \frac{140 \text{ g}}{55.8 \text{ g/mole}} \\
 &= 2.51 \text{ moles}
 \end{aligned}$$

According to the above equation, for the oxidation of 4 moles of iron into iron (III) oxide 3 moles of oxygen are required.

$$\begin{aligned}
 \text{No. of moles of oxygen required to oxidize} &= 2.51 \times \frac{3}{4} \\
 &= 1.88 \text{ moles.}
 \end{aligned}$$

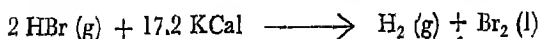
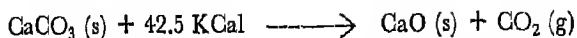
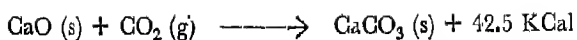
$$\begin{aligned}
 \text{Weight of 1.88 moles of oxygen} &= 1.88 \times 32 \\
 &= 60.2 \text{ g}
 \end{aligned}$$

(b) 4 moles of iron, when oxidized, give 2 moles of iron (III) oxide.

$$\begin{aligned}
 2.61 \text{ moles of iron, when oxidized will give} & \frac{2 \times 2.51}{4} \\
 &= 1.26 \text{ moles}
 \end{aligned}$$

### 5.3 Equations showing the heat of reaction

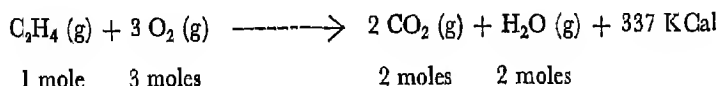
Mention that energy changes accompanying a chemical reaction may be shown by incorporating the quantities of heat absorbed or evolved. Tell the students that if any chemical reaction is endothermic, the reverse reaction is exothermic and involves numerically the same amount of heat change. For example,



Besides the example worked out in the text, the following may also be tried.

### Example 3

The thermochemical equation for the combustion of ethylene gas,  $\text{C}_2\text{H}_4$ , is



- Calculate the amount of heat evolved by burning 14 kg of ethylene gas.
- How many kilograms of water at  $25^\circ\text{C}$  can be heated to  $55^\circ\text{C}$  by the heat evolved in the above reaction?

(Atomic weight of carbon = 12.0, atomic weight of hydrogen = 1.0); specific heat of water = 1.00.

First determine the number of moles of ethylene present in 14 kg i.e., 14000 g of ethylene.

$$\begin{aligned} \text{Number of moles of ethylene} &= \frac{14000 \text{ g}}{28 \text{ g/mole}} \\ &= 500 \text{ moles} \end{aligned}$$

According to the equation, one mole of ethylene, on complete combustion, liberates 337 KCal of heat.

$\therefore$  500 moles of ethylene would liberate  $500 \times 337$  i.e.,  $1.685 \times 10^5 \text{ KCal}$  of heat.

- To heat 1 kg of water from  $25^\circ\text{C}$  to  $55^\circ\text{C}$  the amount of heat required is

amount of water  $\times$  specific heat  $\times$  temperature difference.

$$= 1000 \times 1 \times 30$$

$$= 30 \text{ KCal}$$

30 KCal of heat can raise the temperature of one kilogram of water from 25°C to 55°C.

$$\therefore 168500 \text{ KCal of heat would raise } \frac{168500 \times 1}{30} \\ = 6616.6 \text{ kg of water}$$

#### Example 4

The combustion of 5 g of coke (carbon) raised the temperature one kilogram of water from 20°C to 57°C. Write the thermochemical equation for the combustion of coke.

The amount of heat required to raise the temperature of one kilogram of water from 20°C to 57°C is

$$1 \times 1 \times 37 = 37 \text{ KCal}$$

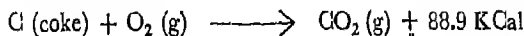
$\therefore$  37 KCal of heat is produced when 5 g of coke is burnt. Number of moles in 5 g of coke

$$= \frac{5 \text{ g}}{12 \text{ g/mole}} = 0.416 \text{ mole}$$

0.416 mole of coke produces 37 KCal of heat

$$\therefore 1 \text{ mole of coke will produce } \frac{37 \times 1}{0.416} = 88.9 \text{ KCal}$$

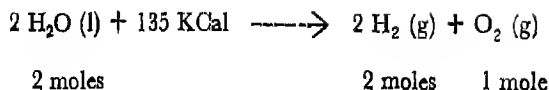
Therefore, the thermochemical equation for the combustion of coke is



#### Example 5

How much heat is required for the complete decomposition of 24 g of water into hydrogen and oxygen ?

The equation for the above reaction is



First determine the number of moles of water present in 24 g of water

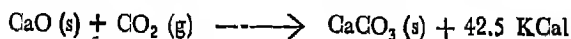
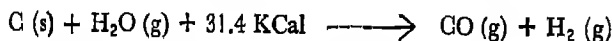
$$\text{No. of moles of water} = \frac{24 \text{ g}}{18 \text{ g/mole}} = 1.33 \text{ moles}$$

According to the equation the amount of heat required for the complete decomposition of 2 moles of water into gaseous hydrogen and oxygen is 136 KCal.

∴ The amount of heat required for the decomposition of 1.33 moles of water

$$= \frac{136 \times 1.33}{2} = 90.5 \text{ KCal}$$

Besides the reactions discussed in the text, the following may be useful while this topic is dealt with :



#### 5.4 Ionic equations

An illustration of ionic reactions and ionic equations can be brought out clearly through experiment 10.

## Experiment 10 : Reactions involving ions

Time required : 40 minutes

For part A of this experiment 0.5 g of each six substances mentioned are to be given to each student. The observations made in this part help the students in guessing which substances will be precipitated from the solution and also in writing ionic equations for the reactions in Part B.

Aqueous solutions (2%) of the eight substances mentioned in the laboratory manual are to be given for Part B. While writing the equation barium sulphate,  $\text{BaSO}_4$ , barium carbonate,  $\text{BaCO}_3$ , precipitated in the reactions are to be shown with arrows pointing downwards ( $\downarrow$ ) or with the subscript (s).

After the experiment is over, discuss the reactions given in the text.

## BACKGROUND INFORMATION

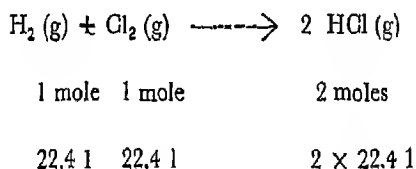
### 5.1 Meaning of an equation

#### Molar volume of a gas

The quantity of a substance whose weight in grams is numerically equal to its molecular weight is called a gram-molecular weight or a mole of that substance. The volume occupied by one mole (1 gram molecular weight) of a gas at STP\* is called its molar volume. It is determined experimentally by weighing a certain volume of the gas under known conditions of temperature and pressure. The volume is then reduced to STP from which the molar volume is calculated.

It has been found that 1 mole of any gas occupies 22.4 litres at STP. Since, according to Avogadro's law, equal volumes of all gases contain equal number of molecules under similar conditions of temperature and pressure, it follows that 1 mole of any gas\*\* contains the same number of molecules represented by  $N$  and is equal to  $6.023 \times 10^{23}$  (Avogadro's number)

Extending the above information we may further interpret the following equation already dealt with in the text.



The equation may be read as : 22.4 litres of hydrogen react with 22.4 litres of chlorine to yield  $2 \times 22.4$  litres of hydrogen chloride (all the volumes being at STP).

---

\* STP = Standard Temperature and Pressure

Standard Pressure = 760 mm Hg

Standard Temperature =  $0^\circ\text{C}$  or  $273 \text{ K}$

\*\* It also applies to 1 mole of a solid or liquid.

### 5.3 Equations showing the heat of reaction

Every chemical substance possesses energy which is closely related to its nature and composition. This energy, which may be said to be the energy stored during its formation, can be broadly classified into two kinds – the kinetic energy due to the different types of motions of the molecule and the potential energy due to attractive and repulsive interactions between molecules and molecules and between atoms within a molecule.

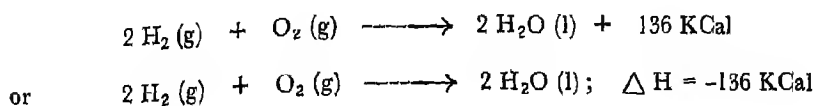
It is a matter of considerable difficulty to calculate the heat content of a substance. Further, these are not of great importance in thermochemical equations. The heat effect in a chemical reaction, measures the difference between the total heat content of the products and the total heat content of the reactants. It is only the difference that is of interest to a chemist. If the total heat content of the reactants is more than that of the products heat is released during the reaction. Conversely if the total heat content of the products is more than that of the reactants heat is absorbed during the reaction.

The heat change in chemical reactions is indicated by " $\Delta H$ "; the Greek letter  $\Delta$  (delta) indicates the difference.

$$\Delta H = \left| \begin{array}{c} \text{Total heat content} \\ \text{of products (H}_p\text{)} \end{array} \right| - \left| \begin{array}{c} \text{Total heat content} \\ \text{of reactants (H}_r\text{)} \end{array} \right|$$

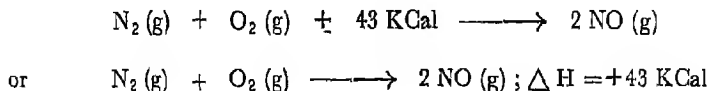
$\Delta H$  may be positive or negative. The sign  $\Delta H$  of has a definite meaning. It is positive for endothermic reactions and negative for exothermic ones. Since  $\Delta H$  varies with temperature it is necessary that mention should be made of the temperature at which the reaction is carried out. Most often the  $\Delta H$  values given in literature refer to 25°C unless otherwise specified.

The following is an example of the two ways of writing the same equation to express the heat change.



Both equations indicate that when 4 g of gaseous hydrogen react with 32 g of gaseous oxygen to form 36 g of liquid water, 136 KCal of heat is released. Since 136 KCal are released to the surroundings, the plus (+) sign is used in the first equation. Since the total heat content of the product is less than that of the reactants, a loss in heat content has taken place and we indicate this fact by using a minus (-) sign before the  $\Delta H$  value.

Similarly, we can write two equations to express the heat change for the formation of gaseous nitrogen monoxide from gaseous nitrogen and oxygen.



Both equations indicate that 43 KCal of heat is absorbed when 28 g of gaseous nitrogen react with 32 g of gaseous oxygen to form 60 g of gaseous nitrogen monoxide. Since 43 KCal are absorbed, the heat change is shown on the left hand side of the first equation. Since the heat content of the product is more than the heat content of the reactants a rise in heat content has taken place, and we indicate this fact by using a plus (+) sign with the  $\Delta H$  value.

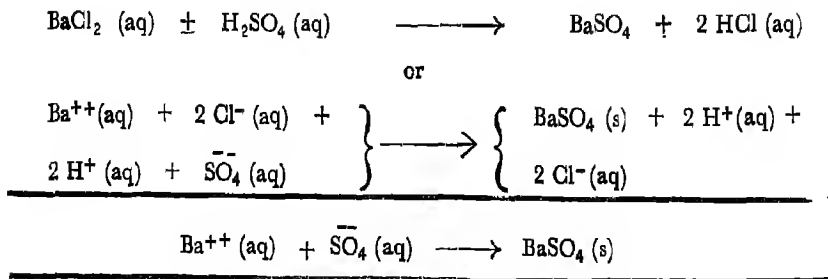
#### 5.4 Ionic equations

The following procedure is usually adopted for balancing ionic equations.

1. Write the balanced equation.
2. Write the same equation in terms of ions.
3. Cancel the common ions on both sides of the equations to get the net ionic equation.

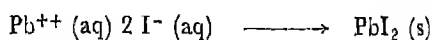
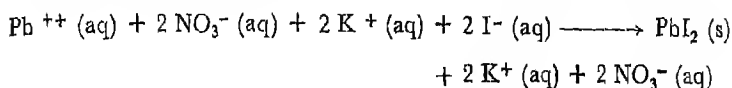
This is illustrated by the reactions between aqueous solutions of

- (i) barium chloride,  $\text{BaCl}_2$ , and sulphuric acid,  $\text{H}_2\text{SO}_4$
  - (ii) lead nitrate,  $\text{Pb}(\text{NO}_3)_2$ , and potassium iodide, KI
  - (iii) ferric chloride,  $\text{FeCl}_3$ , and ammonium hydroxide,  $\text{NH}_4\text{OH}$ .
- (i) The complete equation for the reaction is:

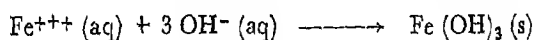
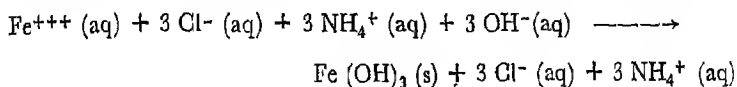
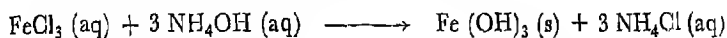




(ii) The complete equation for the reaction is :



(iii) The complete equation for the reaction is :

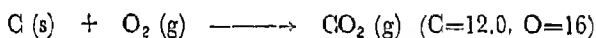


## Quiz

Each question is followed by four answers. One of them is correct. Choose the correct answer and indicate in the bracket provided.

1. A chemical equation gives ( )
- (a) information about the temperature at which the reaction takes place
  - (b) information about the pressure at which the reaction takes place
  - (c) information about the rate at which the reaction proceeds
  - (d) the quantities of reactants and of products involved and their states of aggregation.

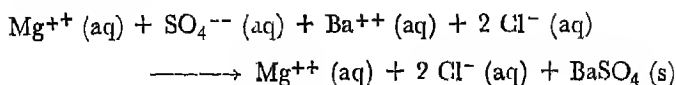
2. Consider the following reaction .



- (i) The number of moles of carbon dioxide formed from 2.40 g of carbon is ( )
- (a) 2.40 moles
  - (b) 1.00 mole
  - (c) 1.20 moles
  - (d) 0.20 mole
- (ii) The number of grams of carbon dioxide formed from 2.40 g of carbon is ( )
- (a) 8.8 g
  - (b) 5.6 g
  - (c) 4.4 g
  - (d) 2.4 g
- (iii) The volume of carbon dioxide in litres at S.T.P. obtainable from 2.4 g of carbon is ( )
- (a) 22.4 litres
  - (b) 11.2 litres
  - (c) 4.48 litres
  - (d) 2.24 litres
3. 44 g of carbon dioxide is produced by heating 1 mole of calcium carbonate. To produce 2.2 g of  $\text{CO}_2$  the number of moles of calcium carbonate required would be ( )
- (a) 0.4
  - (b) 0.2
  - (c) 0.1
  - (d) 0.05

4. When 18 grams of water is decomposed, the products are 16 g of oxygen and ( )
- 16 g of hydrogen
  - 4 g of hydrogen
  - 2 g of hydrogen
  - 1 g of hydrogen
5. Potassium] chlorate on heating decomposes to give oxygen according to the equation
- $$2 \text{KClO}_3 \longrightarrow 2 \text{KCl} + 3 \text{O}_2$$
- The volume of oxygen liberated at S.T.P. by heating 1 mole of  $\text{KClO}_3$  is ( )
- 22.4 litres
  - 33.6 litres
  - 44.8 litres
  - 11.2 litres
6. An example of an endothermic reaction is ( )
- burning of carbon in oxygen
  - burning of hydrogen in oxygen to produce water
  - decomposition of water to gaseous hydrogen and oxygen
  - formation of slaked lime from quick lime.
7. Hydrogen and oxygen react according to the following reaction ( )
- $$\text{H}_2 (\text{g}) + \frac{1}{2} \text{O}_2 (\text{g}) \longrightarrow \text{H}_2\text{O} (\text{l}) + 68.0 \text{ kcal}$$
- The heat liberated when 4 moles of gaseous hydrogen reacts with 2 moles of gaseous oxygen to form liquid water is ( )
- 63 kcal
  - 136 kcal
  - 204 kcal
  - 272 kcal
8. In the precipitation reaction between sodium sulphate solution and barium chloride solution the actual ionic reaction takes place between ( )
- sodium ions and chloride ions
  - barium ions and sulphate ions
  - sodium ions and sulphate ions
  - barium ions and chloride ions.

9. An ionic reaction takes place in aqueous solution according to the equation ( )

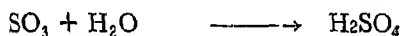
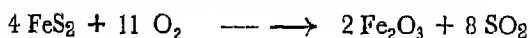


The spectator ions in the reaction are ( )

- (a)  $\text{Mg}^{++}$  and  $\text{SO}_4^{--}$  (b)  $\text{Mg}^{++}$  and  $\text{Ba}^{++}$   
 (c)  $\text{Mg}^{++}$  and  $\text{Cl}^{-}$  (d)  $\text{Ba}^{++}$  and  $\text{SO}_4^{--}$

10. Calculate the weight of lime (calcium oxide) that can be obtained by heating 100 kg of limestone that is 94% pure  $\text{CaCO}_3$ .

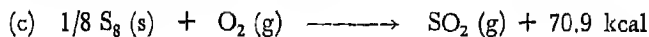
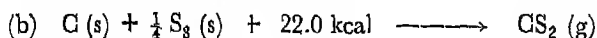
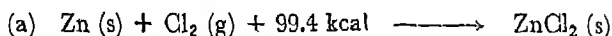
11. How many Kg of pure  $\text{H}_2\text{SO}_4$  is obtained from 10 Kg of pure iron pyrites ( $\text{FeS}_2$ ) according to the following equations ?



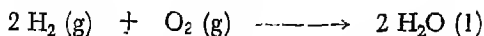
12. A mixture containing 10 g of hydrogen and twice its weight of oxygen are sparked to form water. Calculate the weight of water formed.

13. An element A has an atomic weight of 12. Element B has an atomic weight of 35.5. If one atom of A combines with 4 atoms of B calculate the molecular weight of the compound formed by combination of A and B.

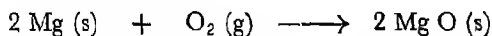
14. Explain the meaning of the following equations :



15. State all the information provided by the equation :

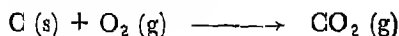


16. What weight of magnesium oxide will be formed when 30.48 g of magnesium react with oxygen ?



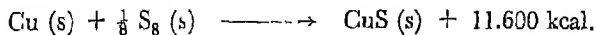
(Atomic weight of magnesium = 24.3, atomic weight of oxygen = 16.0)

17. Calculate the number of grams of oxygen that react with 60 g of carbon to form carbon dioxide.



18. Calculate the number of grams of potassium chloride, KCl, that would be formed along with 10 g of oxygen in the decomposition of potassium chlorate,  $\text{KClO}_3$ .

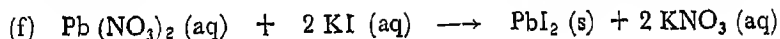
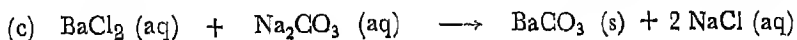
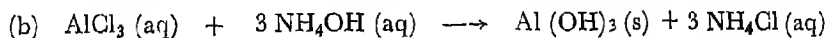
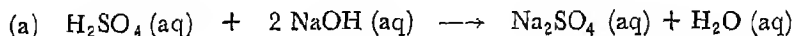
19. Copper metal reacts with sulphur according to the equation.



Calculate the number of calories that would be liberated by the reaction of (a) 10 g of sulphur, (b) 63.54 g of copper.

20. When 12 g of carbon reacts with sulphur to form carbon disulphide,  $\text{CS}_2$ , 22000 cal are absorbed. If 3 g of carbon completely reacts (to form  $\text{CS}_2$ ), calculate the quantity of heat absorbed.

21. Write the net ionic equation for each of the following reactions.



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# Hydrogen, Oxygen and Water

## INSTRUCTION MATERIAL

### Introduction

Hydrogen and oxygen are among the industrially important elements. As is well known oxygen is essential to life. The importance of water to life and industry cannot be over emphasised. This chapter deals with these important aspects of these substances—hydrogen, oxygen and water.

### Outline

1. Hydrogen (6.1)
2. Oxygen (6.2)
3. Oxidation and reduction (6.3)
4. Ozone (6.4)
5. Water (6.5)
6. Hardness of water (6.6)
7. Composition of water (6.7)
8. Properties of water (6.8)
9. Hydrogen peroxide (6.9)

## Concepts

1. Oxidation is a process involving a loss of electrons; reduction involves a gain of electrons.
2. Polarity of the water molecule is a major factor responsible for the solvent action of water.
- 3 The existence of an element in more than one form in the same physical state is known as *allotropy*.

## SCHEDULE

Lesson	Experiment	Postlab	Demonstration	Text covered
1	—	—	—	6.1 and 6.2
2	—	—	—	6.3 and 6.4
3	—	—	—	6.5
4	11, 12	Discussion	—	6.6
5	13	Discussion	—	6.6
6	—	—	1	6.7
7	14	—	2	6.8
8	15	—	—	6.9
9	—	—	—	Review & quiz

## Development

### 6.1 Hydrogen

Elicit from the students the chemistry of hydrogen from what they learnt in the lower classes. Mention the occurrence of hydrogen in the universe and some commercial uses of the gas to emphasise the importance of the element. While it is not required for breathing, hydrogen is nevertheless essential to life, because of its presence in water and in many of the compounds that make up plant and animal tissue. It is important as a component of many fuels – gasoline, kerosene, fuel oil and natural gas – and as such plays an important role in industry. Besides it is an important raw material in the manufacture of many chemical substances like ammonia. It is also used in the hydrogenation of oils. Point out why the laboratory method of preparing the gas is not suitable for large scale manufacture (especially the cost and economic feasibility of the process). While describing any industrial method stress the main principles involved. Ask the students to write the chemical equations for the reactions involved.

## 6.2 Oxygen

The sequence for discussion is the same as in the case of hydrogen. Introduce the phenomenon of allotropy by first recalling the existence of carbon and sulphur in different forms about which the students learnt in the lower classes (the word allotropy was not used earlier). Point out that allotropy is a phenomenon exhibited by many other elements besides carbon and sulphur, for example, oxygen, phosphorus, tin etc.

One of the widely used methods for the industrial preparation of oxygen is *fractional distillation*, of liquid air. When air (previously freed from carbon dioxide, moisture, etc.) is compressed to about 200 atmospheres and is then suddenly released into a region of low pressure or a vacuum, there is a considerable fall in temperature due to the expansion. The cooled air is recompressed and then released again into the vacuum when further cooling is produced. This cycle of operations is repeated until the air is cooled to such an extent that it liquefies. This liquid, which consists mainly of oxygen and nitrogen (and a minute quantity of a mixture of the noble gases argon, neon, krypton and xenon) is now subjected to fractional distillation. This operation makes use of the fact that liquid nitrogen boils at a lower temperature ( $-196^{\circ}\text{C}$ ) than liquid oxygen ( $-183^{\circ}\text{C}$ ). When the liquid air is passed through a specially constructed tall tower called a *fractionating column* (Fig 6.1) the vapours contain a

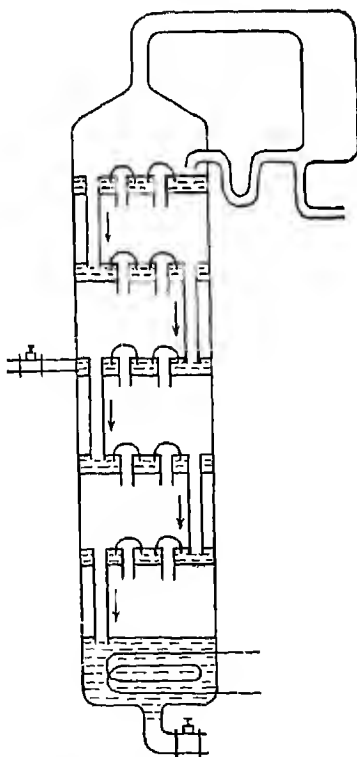


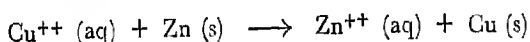
Fig. 6.1 The fractionating column



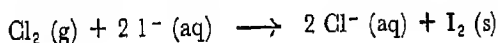
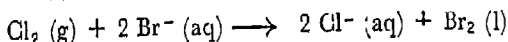
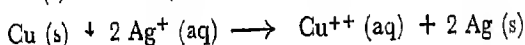
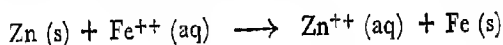
greater proportion of the more volatile nitrogen, while the residual liquid contains a greater proportion of oxygen. By repeating this process a vapour which is almost pure nitrogen and a residual liquid of almost pure oxygen can be obtained.

### 6.3 Oxidation and reduction

The electronic concept of oxidation and reduction should be introduced only formally and no attempt need be made at this stage to explain all oxidation-reduction reactions on the basis of loss or gain of electrons. A few more simple examples of oxidation and reduction reactions in addition to those mentioned in the text may be given. A good example is that of the displacement of copper from aqueous solutions of copper salts by zinc metal. In this reaction copper ions gain two electrons and hence, are reduced. While zinc atoms lose two electrons and are oxidized. The reaction can be represented as:



The *Copper (II) ion*,  $\text{Cu}^{++}$  is the *oxidizing agent* as it oxidizes zinc to the zinc ion,  $\text{Zn}^{++}$ . *Metallic zinc*, in turn, is the *reducing agent* as it reduces copper (II) ion to metallic copper. The oxidizing agent gets reduced, while the reducing agent gets oxidized. Emphasize this point and the fact that oxidation and reduction occur simultaneously. Write out a few equations on the black board and ask the students to point out the oxidizing agent and reducing agents. Some examples are given below.



### 6.4 Ozone

At this stage no more than what is given in the text need be discussed.

### 6.5 Water

Importance of water for all plant and animal life could be emphasized through common examples. It should be interesting to talk here about the problem relating to quest for water. It would be relevant to mention the types of natural waters and the nature of impurities associated with each type. The methods described in the text for the purification of water for drinking purposes should be explained.

### 6.6 Hardness of water

Hardness of water and its removal may be introduced through experiments 11, 12 and 13. The postlab discussion should bring out the ideas given in the text.

## Experiment 11 : Detection of hardness

Time required : 30 minutes

1. Prepare the soap solution required for this experiment by dissolving 2 g of any good quality soap (not soapless soap) in 100 ml distilled water. Alternately commercially available soap solution can be used. Dilute if necessary and use.
2. In place of sea water an aqueous solution containing about 2 g sodium chloride; 1 g magnesium sulphate, and 0.1 g calcium chloride per 100 ml water may be provided
3. If well water is not available, water from any other source may be provided.

The number of drops of soap solution required to give a permanent lather with any sample of water depends on a number of factors such as the quality of the soap itself, the nature of the water samples collected. The details given here for preparation of different solutions serve only as general guidelines. Alter the concentration according to requirements and *standardise the results by doing the tests yourself before assigning this experiment to the students.*

### Postlab discussion

After completion of the experiment collect the data obtained by students and arrange the various samples of water in order of increasing hardness; the order would be the same as in the increasing order of the number of drops of soap solution required to give a permanent lather with water (permanent here may be taken to mean lather lasting atleast for 2 minutes).

## Experiment 12 : Causes of hardness of water

Time required : 30 minutes

1. Soap solution, as in experiment 11.
2. The various other solutions required for this experiment (excepting solution of calcium bicarbonate) are to be prepared by dissolving about 1 g of the salt in about 4 litres of ordinary water.
3. Aqueous solution of calcium bicarbonate is to be prepared by passing carbon dioxide (from a Kipp's apparatus) through a suspension of calcium carbonate (1 g in 4 litres) until a clear solution is obtained. Or pass carbon-dioxide through freshly prepared lime water until a clear solution is obtained.

### **Postlab discussion**

Ask the students to record the information obtained during the experiment. Elicit from the students the answers to the questions given in laboratory manual. The discussion would enable the students to understand the causes of hardness of water.

### **Question and Answer**

Based on your observations,

- (a) Which ions present in the different solutions cause hardness of water?
- (b) Which ions do not cause hardness?

- (a) Calcium and magnesium ions,
- (b) Sodium ions.

### **Experiment 13 : Removal of hardness in water**

Time required : 30 minutes

1. Soap solution, as in experiment 11
2. Aqueous solutions of calcium bicarbonate, calcium chloride, and magnesium sulphate, as in experiment 12.
3. Sodium carbonate solution is to be prepared by dissolving about 1 g of the carbonate in 2 litres water

### **Postlab discussion**

The results obtained by students should be discussed in accordance with the ideas presented in section 6.6 of the text regarding removal of hardness. The discussion should lead students to an understanding of the disadvantages of hardness and of the method used for its removal.

### **Questions and answers**

1. In which of the solutions used, is hardness removed by boiling? In which is it not?

The hardness is removed by boiling in the case of the solution containing calcium bicarbonate. It is not removed in the case of the other solutions.

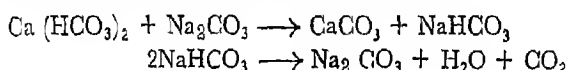
2. Which of the solutions need chemical treatment for the removal of hardness? The solutions containing calcium chloride and magnesium sulphate.

Although the students have performed the experiments dealing with the causes and removal of hardness of water, nothing has been said about its estimation. The determination of hardness of water may be introduced through Demonstration I. Explain

how natural waters may contain bicarbonates of calcium and magnesium which cause temporary hardness, and chlorides and sulphates of calcium and magnesium which cause permanent hardness. It is generally expressed in parts of calcium carbonate per million parts of water by weight (which is the same as milligrams per litre of water).

Temporary hardness can be estimated by titrating bicarbonates with a standard dilute acid, say 0.05 M HCl, using methyl orange as indicator. To estimate permanent hardness, the water is treated with sodium carbonate and after complete precipitation of the alkaline earth salts, the solution is filtered and the excess of carbonate determined by titrating it with a standard acid (0.05 M HCl). The amount of sodium carbonate used up is then calculated.

Calcium sulphate which causes permanent hardness reacts with  $\text{Na}_2\text{CO}_3$  forming  $\text{CaCO}_3$  and  $\text{Na}_2\text{SO}_4$ , the former of which is removed by filtration. Calcium bicarbonate which is responsible for temporary hardness also reacts with  $\text{Na}_2\text{CO}_3$  added but it does not interfere with the determination of permanent hardness, as an equivalent amount of sodium carbonate is regenerated according to the reactions



### Demonstration I Determination of hardness of water

Time required 50 minutes.

#### Materials required

Burette	Hydrochloric acid 0.05 M	1 litre
Burette stand with clamps	Sodium carbonate 0.25 M	1 litre
Porcelain tile	Methyl orange (indicator)	
Conical flask, 250 ml	Sample of hard water	
Evaporating dish, 500 ml		
Glass rod		
Pipette (25 ml)		
Water bath		

#### (a) Preparation of reagents

1. Concentrated hydrochloric acid is approximately 10 M. To prepare 1 litre of 0.05 M HCl, dissolve 5 ml of concentrated HCl in distilled water and make the volume to 1 litre. Find out the exact molarity of this solution by titration against standard sodium carbonate solution.
2. To prepare 0.025  $\text{Na}_2\text{CO}_3$ , dissolve 2.65 g of anhydrous sodium carbonate in water, and make the solution to 1 litre.
3. A 0.1% solution of methyl orange indicator is required. To prepare this, dissolve 10 mg of the solid indicator in 10 ml of 70% ethanol.

4. Tap water from the laboratory may be treated as a sample of hard water.

(b) **Determination of temporary hardness**

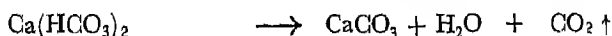
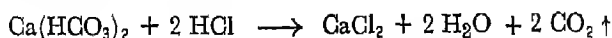
Fill a clean burette with 0.05 M HCl solution. Pipette out 100 ml of the given sample of hard water into a conical flask. Add 2 or 3 drops of methyl orange indicator. The colour of the solution will be straw yellow. Run in the acid solution from the burette drop-wise, with constant swirling, until the colour of the solution just turns pink. Note the volume of the acid solution required for this colour change. Enter the results in a tabular form given below. Repeat the experiment to obtain a set of concordant readings. Let the mean volume of acid solution used in the titration be  $V_1$  ml.

BURETTE READINGS		
Initial	Final	Vol. of 0.05 M HCl used (ml)

**Calculations**

$$V_1 \text{ ml of } 0.05 \text{ M HCl} = \frac{(V_1 \times 0.05)}{1000} \text{ moles of HCl}$$

The following reactions are involved



The above equations indicate :

2 moles of HCl reacted  $\equiv$  1 mole of  $\text{CaCO}_3$  present

$$\therefore \frac{(V_1 \times 0.05)}{(1000)} \text{ moles of HCl} \equiv \frac{1}{2} \frac{(V_1 \times 0.05)}{(1000)} \text{ moles of } \text{CaCO}_3$$

This  $\text{CaCO}_3$  is present in 100 ml of hard water; the amount of  $\text{CaCO}_3$  present in one million parts, therefore, is

$$\begin{aligned}
 &= \frac{1}{2} \times \frac{(v_1 \times 0.05)}{(1000)} \times \frac{1,000,000}{100} \\
 &= (0.25 \times v_1) \text{ moles} \\
 &= (25 \times v_1) \text{ parts / million parts water, since molecular weight of } \text{CaCO}_3 \text{ is } 100. \text{ Let this be } w_1.
 \end{aligned}$$

### (c) Determination of permanent hardness

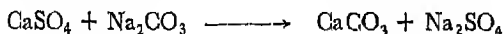
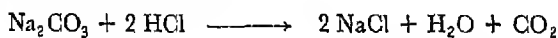
#### Procedure

Fill a clean burette with 0.05 M hydrochloric acid solution. Pipette out 50 ml of 0.025 M sodium carbonate solution. Pipette out 50 ml of 0.025 M sodium carbonate solution into a conical flask. Add 2 or 3 drops of methyl orange indicator. The colour of the solution will be straw yellow. Run in the acid solution from the burette until the colour changes to pink (keep the solution in the conical flask swirling during the titration). Note the volume of the acid solution added. Let it be A ml. Tabulate the readings as before. (This titration will not be necessary if the strength of the sodium carbonate solution is known accurately).

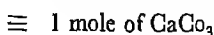
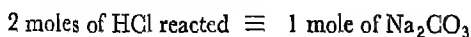
(ii) Pipette out 100 ml of the given sample of hard water into a 500 ml evaporating dish. Add exactly 50 ml of 0.025 M sodium carbonate. Evaporate the solution to dryness on a water bath. See that there is no spurting. Dissolve the residue in freshly prepared distilled water. Filter off any undissolved substance. Collect the filtrate in a 250 ml conical flask. Wash the filter paper thoroughly with small quantities of distilled water and collect the washings also in the conical flask. Add 2 or 3 drops of methyl orange indicator. The colour of the solution will be yellow. Run in the acid solution from the burette till the colour of the solution changes to pink. Note the volume of the acid required for the titration. Let it be B ml

#### Calculations

The following reactions are involved :



The above equation indicate that



50 ml of 0.025 M  $\text{Na}_2\text{CO}_3$  requires A ml of 0.05 M HCl, and after the reaction 50 ml of  $\text{Na}_2\text{CO}_3$  requires 'B' ml of the HCl.

$$= v_2 \text{ ml (say)}$$

$$v_2 \text{ ml of } 0.05 \text{ M HCl} \equiv$$

$$\frac{(v_2 \times 0.05)}{1000} \text{ moles of HCl}$$

Since 2 moles of HCl reacted  $\equiv$  1 mole of  $\text{CaCO}_3$  present

$$\frac{(v_2 \times 0.05)}{1000} \text{ moles of HCl reacted} \equiv \frac{1}{2} \frac{(v_2 \times 0.05)}{1000} \text{ moles of CaCO}_3$$

This  $\text{CaCO}_3$  is present in 100 ml of water.

$\therefore$  amount present in one million parts water

$$= \frac{1}{2} \frac{(v_2 \times 0.05)}{1000} \times \frac{1,000,000}{100} = (0.25 \times v_2) \text{ moles}$$

$$\equiv (25 \times v_2) \text{ g/million g water}$$

Let this be  $w_2$

$$= \text{Temporary hardness} + \text{Permanent hardness}$$

$$= (w_1 + w_2) \text{ p.p.m.}$$

## 6.7 Composition of water

The composition of water by volume can be verified by analysis. Experiment 14 fits in very well here.

### Experiment 14: Composition of water

Time required : 30 minutes

For this experiment about 1% aqueous solution of sodium sulphate may be used.

### Question and answer

What is the volume ratio of the gases liberated in this experiment? Is this ratio fixed or variable?

The volume ratio of hydrogen to oxygen is 2:1. This ratio is fixed.

## 6.8 Properties of water

Ideas developed in chapter 4 '*How atoms combine*' should be recalled in discussing the polar nature of water. The solvent action of water can then be explained on the basis of polarity of the water molecule.

The phenomena of deliquescence and efflorescence can be introduced by exhibiting to students crystals of calcium chloride and sodium thiosulphate previously exposed to air. Ask the students to point out the two different behaviours.

The action of metals on steam can be introduced through Demonstration II

### Demonstration II : Action of metals on steam

Time required : 40 minutes

#### Materials required

Hard glass tube, 15 cm x 2.5 cm

Distilled water

Round bottomed flask, 250 ml

Magnesium ribbon 20 cm

Burners

Iron powder 10 g

Stopper (two holed)

Stoppers (one holed) 2

Bent tubes 2

Glass tube (0.5 x cm x 30)

Splinter

Trough

Gas jar

Beehive shelf

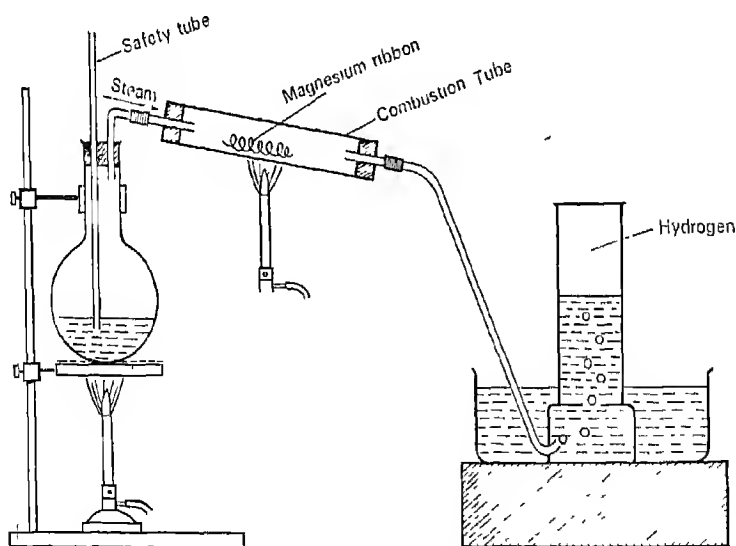


Fig. 6.2 Action of steam on metals



## Procedure

Set up the apparatus as shown in Fig. 6.2. Rub a piece of magnesium ribbon (about 15 cm) with sand paper until it is bright and shiny. Place it in the form of a coil in the middle of the hard glass tube.

Boil the water gently. Warm the glass tube so that steam does not condense when it enters the hard glass tube. When all the air has gone out of the tube, invert the gas jar filled with water over beehive shelf kept in a trough filled with water.

Heat the magnesium ribbon and pass the steam slowly. Let the students observe what happens to the magnesium and in the gas jar. Test the gas collected in the jar with a burning splinter to show that the gas collected is hydrogen. Explain the reaction involved.

Cool the glass tube after removing the delivery tube, clean and dry it. Repeat the experiment using iron powder in place of magnesium.

## 6.9 Hydrogen peroxide

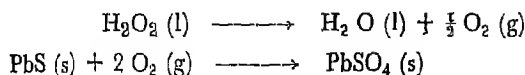
Mention the importance of this substance as an oxidizing agent. Explain its uses. Important reactions of hydrogen peroxide can be introduced through experiment 15.

### Experiment 15 : Reactions of hydrogen peroxide

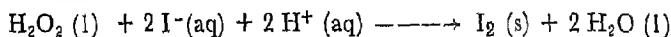
Time required : 40 minutes.

1. Dilute sulphuric acid is to be prepared by diluting the concentrated acid 10 times. (Caution : acid should be added slowly to water and not water to acid).
2. Other solutions (excepting potassium permanganate solution) should be about 2% strong.
3. Potassium permanganate solution should be about 0.1% strong.
4. Hydrogen peroxide solution is to be prepared by diluting the commercially available (10 volume) solution ten times.
5. A Kipp's generator may be used for preparing hydrogen sulphide. As an alternative, a mixture of powdered sulphur with scraped paraffine wax (1:3) may be heated in a test tube fitted with a delivery tube to generate hydrogen sulphide. The reaction ceases as soon as heating is stopped.

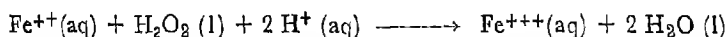
The black precipitate (PbS) turns white due to the formation of lead sulphate (PbSO<sub>4</sub>).



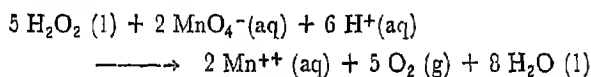
When hydrogen peroxide is added to acidified potassium iodide solution, a dark brown precipitate of iodine is formed.



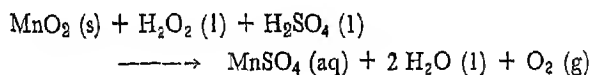
When hydrogen peroxide is added to an acidified solution containing iron (II) ions (colourless or pale green), the solution turns yellow due to the formation of iron (III) ions.



When hydrogen peroxide is added to acidified potassium permanganate solution, the latter is decolourised.



When manganese dioxide is added to hydrogen peroxide there is a brisk evolution of oxygen gas. If a glowing splinter is introduced into the test tube the splinter bursts into flame.

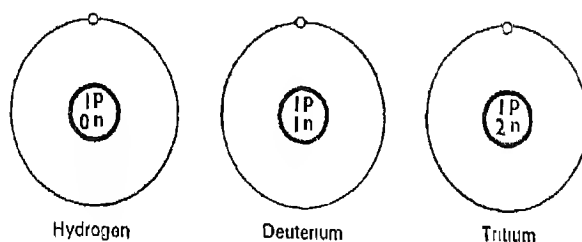


## BACKGROUND INFORMATION

### 6.1 Hydrogen

#### *Isotopes of hydrogen*

Three isotopes of hydrogen are known. Their structures are given below



*Fig. 6.3 Isotopes of hydrogen*

Ordinary hydrogen or protium,  ${}^1_1\text{H}$ , vastly predominates in a sample of naturally occurring hydrogen.

Deuterium,  ${}^2_1\text{H}$ , denoted by the symbol D, is present in ordinary hydrogen to the extent of one part in about six thousand parts. It can be separated from protium by the method of gaseous diffusion or it can be made from deuterium oxide,  $\text{D}_2\text{O}$ , which is commonly known as *heavy water*. The latter is obtained by the electrolysis of dilute aqueous solutions of acid when the hydrogen ions are preferentially discharged at the platinum cathode. By repeatedly adding fresh water and electrolysing, it is eventually possible to produce deuterium oxide of any desired concentration.

Heavy water is important as a moderator in atomic reactors and is widely used for this purpose in atomic powered submarines and surface vessels. Its physical properties are compared with those of water in Table 6.1.

**Table 6.1 Properties of water and heavy water**

Property	Water	Deuterium oxide
Melting point, °C	0	3.8
Boiling point, °C	100	101.4
Density, g/ml	1.00	1.11
Maximum density at	4°C	11.6°C

The third isotope of hydrogen, which is known as tritium,  ${}^3_1\text{H}$ , is present only in traces in ordinary hydrogen, but it can be made artificially. However, it is radioactive and decays by  $\beta$  emission (to helium) with a half-life of about 12 years.

The isotopes of hydrogen differ noticeably in their physical properties owing to the large percentage difference in their masses, but they are very similar chemically. Deuterium is slightly less reactive than hydrogen.

## 6.2 Oxygen

The fact that oxygen exists in two molecular forms has already been mentioned, and the properties of the ordinary form, consisting of  $\text{O}_2$  molecules, have been discussed. We have also considered the properties of the less common variety—ozone, the molecular formula of which is  $\text{O}_3$ . Different forms of the same element existing in the same physical state are known as *allotropic modifications*, or simply as *allotropes*. Allotropy results from the possibility that the atoms or molecules of an element can sometimes be arranged so as to attain reasonably stable configurations in more than one way. It is exhibited by a number of elements, especially nonmetals.

Two crystalline allotropes of carbon are known, diamond and graphite. In addition other forms of carbon exist, known collectively as amorphous carbons which, in fact, are microcrystalline in structure. Three allotropes of tin—grey tin, white tin and rhombic tin are known. Phosphorus exists in several allotropic forms of which the white (or yellow) and the red forms are better known. Several forms of sulphur are also known.

## 6.5 Water

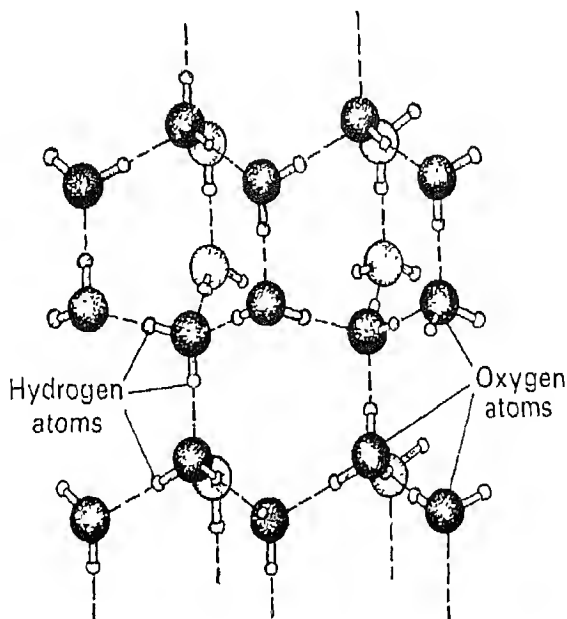
### *The three states of water*

The three physical states of water are steam (gas), water (liquid) and ice (solid). They are chemically identical. Because of its small molecular weight (18) water would be expected to be a gas at ordinary temperature. The fact that it is liquid at ordinary temperatures indicates that water molecules probably form groups of many  $\text{H}_2\text{O}$  units. This will be clear from the following discussion:

It has already been noted that water is a polar molecule because the electrons are attracted more towards the oxygen atom making it negatively charged with respect to hydrogen atoms which become slightly positively charged. Since opposite charges attract, an attractive interaction between the positively charged hydrogen of one water molecule and the negatively charged oxygen of a neighbouring water molecule takes place forming what is called a *hydrogen bond*.

Studies have shown that water molecules in ice are arranged in such a way that each oxygen atom has four hydrogen atoms as its neighbours, two attached by strong

covalent bonds (that is, by shared electrons) and two more by the weaker hydrogen bonds. Such an arrangement leads to a tetrahedral structure having relatively large empty spaces in it (Fig. 6.4).



*Fig. 6.4 Structure of ice*

When ice is melted, the energy supplied in the form of heat breaks some of the hydrogen bonds in ice and the molecules come close together. Thus, water has a greater density than ice and this is why ice floats on water.

In liquid water also, there is considerable hydrogen bonding thus making water an associated liquid. When water is heated, the energy supplied in the form of heat starts breaking the hydrogen bonds. This process continues till, at  $100^{\circ}\text{C}$ , all the hydrogen bonds break down and separate water molecules are formed.

#### *Water as a solvent*

Water is a universal solvent. It dissolves many substances. There are hardly any substances which it will not dissolve, atleast in a minute quantity. It is, therefore, rather difficult to keep it pure. Sea water has almost all the elements found on earth in larger or smaller quantities. Water is relatively inert. This fact is very important biologically, as nutrients can be conveyed, in an unaltered form, to the site where they are needed. Blood plasma, lymph and all other body fluids are essentially aqueous solutions, carrying food to the cells and tissues, or bringing back waste products to various organs from where they can be excreted.

### *Water of hydration – deliquescence and efflorescence*

Many salts combine with water to form hydrates. This phenomenon, known as hydration, results in the formation of solids such as  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ ;  $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$  and  $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ . Sometimes a single salt forms a series of hydrates, as for instance,  $\text{CaCl}_2 \cdot \text{H}_2\text{O}$ ;  $\text{CaCl}_2 \cdot 2\text{H}_2\text{O}$ ,  $\text{CaCl}_2 \cdot 4\text{H}_2\text{O}$ , and  $\text{CaCl}_2 \cdot 6\text{H}_2\text{O}$ . The water molecules are weakly bonded to the metal ions in the salts.

Hydrates have a definite tendency to lose, under certain conditions, their water of hydration. When the vapour pressure of water in a hydrate is greater than the partial pressure of water in the atmosphere, the hydrate tends to lose the water of hydration. It is said to be *efflorescing* or spontaneously dehydrating.

Some anhydrous salts absorb moisture from the atmosphere to form a hydrate, and then continue to absorb moisture till, finally, the hydrate dissolves. A substance that absorbs water vapour till it goes into solution is said to *deliquesce*. The process itself is known as *deliquescence*. Anhydrous calcium chloride is a typical deliquescent substance. It absorbs water vapour from the atmosphere to form the hydrate  $\text{CaCl}_2 \cdot 6\text{H}_2\text{O}$  and then forms the solution. It is often used in the laboratory for removing moisture in desiccators.

Substances which absorb moisture are said to be *hygroscopic*. All deliquescent substances are hygroscopic, but all hygroscopic substances are not deliquescent. Powdered salt, fluor, etc., are hygroscopic to some extent, but they are not deliquescent as they do not absorb enough water to form a solution.

### *Quest for water*

Man has found many ways to remove salt from seawater but the problem is how to do this economically. This aspect has been engaging the attention of authorities throughout the world, especially in areas where water for human and industrial consumption is not easily available.

One of the recent approaches is to produce steam at higher temperature and pressure, use it for generation of electrical power and extract from the power station low temperature steam which after condensation would yield desalinated water.

In effect, large scale desalination will have to be done at dual purpose plants producing both electricity and water. Waste seawater from distillation plants is richer in salts and may be used for their recovery to make water still cheaper. Such plants can, of course, be operated by both conventional and nuclear fuels. The use of nuclear energy, however, seems more advantageous, especially for very large plants. Conventional sources of energy may be more attractive in certain circumstances, but current thinking is mostly in favour of combining desalination units with nuclear units producing electricity. However, research to uncover new desalting processes is still going on all over the world.

## 6.6 Hardness of water

### Ion - exchange resins

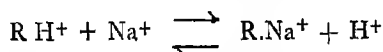
Ion exchange resins are generally organic substances of a complex structure. They are made of chains of carbon atoms or of carbon and nitrogen atoms to which are attached groups such as sulphonic group,  $\text{SO}_3\text{H}$ , or the amino group,  $-\text{NH}_2$ .

There is a considerable cross-linking of the chains such that the whole structure of the resin is a complex three-dimensional net work. These resins are capable of exchanging their ions with similarly charged ions from the solution, and hence the name ion exchange resins. Depending upon the nature of the group attached to the chain, these resins are broadly divided into two types .

- (a) Cation exchange resins contain the groups  $-\text{COOH}$  or  $-\text{SO}_3\text{H}$  and interchange positively charged ions (cations)
- (b) Anion exchange resins contain the group  $-\text{NH}_2$  and interchange negatively charged ions (anions).

The commercial names of some of the resins are DOWEX 50, AMBERLITE IR— 120 (cation exchangers) and DOWEX No. 2, AMBERLITE IR— 4B (anion exchangers).

If a cation exchange resin is immersed in a solution containing cations such as  $\text{Na}^+$  or  $\text{K}^+$ , some of the hydrogens of the acidic groups of the resin will be exchanged for the cations in the solution. Cations such as  $\text{Ca}^{2+}$  or  $\text{Fe}^{3+}$  will require two or three carboxylic groups or sulphonic acid groups on the resin. The interaction between the ions on the resin surface and those in the solution is an equilibrium process and may be represented by the following equation

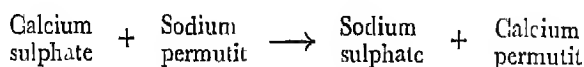


where R represents the surface of the resin ( $\text{RCOO}^-$ ;  $\text{RSO}_3^-$ ) if the resin is stirred with a solution containing  $\text{Na}^+$  ions, an equilibrium is eventually established with as many  $\text{H}^+$  ions being given off by the resin as are taken up by it. Similarly for the  $\text{Na}^+$  ions added to the solution from outside, some of them will be exchanged for more  $\text{H}^+$  ions on the resin until equilibrium is established again. Similarly the  $\text{H}^+$  ions added to the solution will be exchanged in part for  $\text{Na}^+$  ions on the resin.

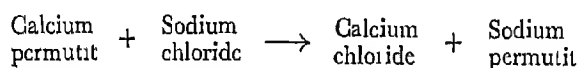
Anion exchange resins are similar to cation exchange resins. They are capable of exchanging  $\text{OH}^-$  or  $\text{Cl}^-$  ions present on their surface for any other anions in the surrounding solutions. If an anion exchange resin is left in contact with a solution, an equilibrium will establish itself just as in the case of cation exchange resins.

(a) *Softening hard water by ion exchange resins.*

When hard water, whether temporary or permanent, is passed over an ion exchange resin, calcium and magnesium ions take their place on the resin while sodium ions from the resin go into solution. *Permutit* is a well known naturally occurring ion-exchange resin used industrially and in the laboratory.



In this method softening is most conveniently done by passing hard water through a column filled with the resin. Once the resin has given up its supply of sodium ions, it cannot soften water further. However, it can be regenerated by percolating concentrated solution of sodium chloride through the resin whereby the above mentioned reaction is reversed.



In this way, the same resin can be used again and again.

(b) *Other methods of softening hard water*

One of the newer methods of softening water depends on the use of *sequestering agents* (agents which tie up metal ions and prevent their precipitation while the water is being used). Two such agents are *sodium hexametaphosphate*, ( $\text{Na}_6\text{P}_6\text{O}_{18}$ ) and *ethylenediaminetetraacetic acid*, EDTA. A preparation using ( $\text{Na}_6\text{P}_6\text{O}_{18}$ ) is sold under the trade name of calgon.



## Quiz

Each question is followed by four answers. One of them is correct. Choose the correct answer and indicate it in the brackets provided.

1. The gas used in the hardening of oils is ( )  
(a) oxygen (b) nitrogen  
(c) hydrogen (d) carbon dioxide.
2. Water gas is obtained by passing steam over ( )  
(a) red hot iron (b) heated aluminium  
(c) heated copper (d) red hot coke.
3. Water gas is used in the manufacture of ( )  
(a) carbon monoxide (b) hydrogen  
(c) oxygen (d) nitrogen.
4. In the manufacture of hydrogen from water gas the catalyst used is ( )  
(a) manganese dioxide (b) finely divided nickel  
(c) sand (d) iron (III) oxide mixed with chromium (III) oxide.
5. Consider the following reactions for the preparation of hydrogen. ( )  
(a)  $3 \text{Fe(s)} + 4 \text{H}_2\text{O(g)} \longrightarrow \text{Fe}_3\text{O}_4\text{(s)} + \text{H}_2\text{(g)}$   
(b)  $2 \text{H}_2\text{O(g)} + \text{C(s)} \longrightarrow 2 \text{H}_2\text{(g)} + \text{CO}_2\text{(g)}$   
(c)  $\text{Zn(s)} + 2 \text{HCl(aq)} \longrightarrow \text{ZnCl}_2\text{(aq)} + \text{H}_2\text{(g)}$   
(d)  $2 \text{Na(s)} + 2 \text{HOH(l)} \longrightarrow 2 \text{NaOH(aq)} + \text{H}_2\text{(g)}$   
(i) The most important industrial preparation of hydrogen gas represented in one of the above reactions is ( )  
(a) reaction (a) (b) reaction (b)  
(c) reaction (c) (d) reaction (d).

- (ii) The reaction which represents the general method of preparing hydrogen in the laboratory is ( )
- (a) a (b) b  
(c) c (d) d.
- (iii) The reaction in which salt is also produced in addition to hydrogen is ( )
- (a) a (b) b  
(c) c (d) d.
- (iv) The reactions in which an oxide is produced in addition to hydrogen gas are
- (a) a and d (b) b and c  
(c) a and c (d) a and b.
6. In the laboratory method of preparation of hydrogen, an acid reacts with ( )
- (a) the metal, iron or zinc  
(b) the metal, sodium or potassium  
(c) the halogen, chlorine or bromine  
(d) the metal, copper or silver.
7. Hydrogen and oxygen readily combine ( )
- (a) at room temperature  
(b) only when warmed  
(c) only when exposed to sunlight  
(d) when ignited by a spark.
8. When steam is passed over hot iron the gas obtained is ( )
- (a) hydrogen (b) oxygen  
(c) carbon (d) carbon monoxide.
9. When hydrogen reacts chemically with nonmetals its atoms attain a stable electronic configuration similar to that of ( )
- (a) radon (b) helium  
(c) argon (d) xenon.

10. Hydrogen combines with chlorine explosively to form hydrogen chloride when ( )  
(a) kept in darkness  
(b) exposed to diffuse day light  
(c) cooled  
(d) exposed to bright sun light.
11. Oxygen readily forms compounds with every element except ( )  
(a) the halogens (b) the alkali metals  
(c) the heavy metals (d) the noble gases.
12. The number of valence electrons in oxygen atom is ( )  
(a) 2 (b) 4  
(c) 6 (d) 8.
13. An example of a covalent oxide which is neutral to litmus is ( )  
(a) sulphur dioxide (b) carbon dioxide  
(c) hydrogen chloride (d) water.
14. Sodium is a good reducing agent because ( )  
(a) it is a metal  
(b) it is soft  
(c) it tends to lose an electron readily  
(d) it has a relatively low melting point.
15. In the reaction between magnesium and oxygen to form magnesium oxide  
(a) magnesium oxidizes oxygen to the oxide ion  
(b) oxygen reduces magnesium to magnesium ion  
(c) magnesium and oxygen are both oxidized  
(d) magnesium is oxidized to magnesium ion.

16. An example of a neutral oxide is ( )  
 (a) carbon monoxide (b) sulphur dioxide  
 (c) sodium monoxide (d) carbon dioxide.
17. Consider the reaction  

$$\text{HgO(s)} + \text{H}_2\text{(g)} \longrightarrow \text{Hg(l)} + \text{H}_2\text{O(g)}$$
- (i) The substance that has been reduced is ( )  
 (a) HgO (b) H<sub>2</sub>  
 (c) Hg (d) H<sub>2</sub>O.
- (ii) The substance that has been oxidized is ( )  
 (a) HgO (b) H<sub>2</sub>  
 (c) Hg (d) H<sub>2</sub>O.
18. In the reaction  $2 \text{K(s)} + \text{Cl}_2\text{(g)} \longrightarrow 2 \text{KCl(s)}$  the oxidizing agent is ( )  
 (a) potassium (b) chlorine  
 (c) potassium ion (d) chloride ion.
19. Fe<sup>++</sup> ion on complete reduction gives ( )  
 (a) Fe atom (b) Fe<sup>+</sup> ion  
 (c) Fe<sup>+3</sup> ion (d) Fe<sup>+4</sup> ion.
20. When sodium burns in chlorine, chlorine is ( )  
 (a) oxidized (b) reduced  
 (c) neither oxidized nor reduced (d) acts as a reducing agent.
21. An example of the element which exhibits phenomenon of allotropy is ( )  
 (a) hydrogen (b) nitrogen  
 (c) oxygen (d) sodium.

22. Which of the following *does not* represent a pair of allotropes? ( )
- (a) hydrogen – deuterium
  - (b) oxygen – ozone
  - (c) red phosphorus – white phosphorus
  - (d) rhombic sulphur – monoclinic sulphur.
23. Oxygen and ozone are ( )
- (a) isotopes
  - (b) allotropes
  - (c) reducing agents
  - (d) combustible substances.
24. Ozone turns ( )
- (a) red litmus paper blue
  - (b) potassium dichromate paper green
  - (c) starch iodide paper blue
  - (d) starch paper blue
25. Sterilization of water is done by adding ( )
- (a) bleaching powder
  - (b) sodium carbonate
  - (c) aluminium sulphate
  - (d) sodium bicarbonate.
26. Hardness of water is caused by ( )
- (a)  $\text{Na}^+$  ions
  - (b)  $\text{K}^+$  ions
  - (c)  $\text{Ca}^{++}$  and  $\text{Mg}^{++}$  ions
  - (d) any metal ions.
27. Temporary hardness of water is removed by adding ( )
- (a) slaked lime in excess
  - (b) slaked lime in requisite amount
  - (c) sodium bicarbonate
  - (d) magnesium bicarbonate.

28. Permanent hardness of water is removed by ( )
- (a) boiling
  - (b) adding slaked lime
  - (c) adding washing soda
  - (d) adding baking soda.
29. Water contains hydrogen and oxygen in the volume ratio of ( )
- (a) 1 : 1
  - (b) 1 : 2
  - (c) 2 : 1
  - (d) 2 : 3
30. An example of the reducing action of  $\text{H}_2\text{O}_2$  is ( )
- (a) liberation of  $\text{I}_2$  from acidified KI solution
  - (b) decolourisation of acidified  $\text{KMnO}_4$  to liberate  $\text{O}_2$
  - (c) converting iron (II) salts into iron (III) salts
  - (d) changing black lead sulphide to white lead sulphate.

## Some Chemical Families

### INSTRUCTION MATERIAL

#### Introduction

The purpose of this chapter is to show that elements with similar properties can be grouped to form a chemical family. This similarity in properties of elements in a family can be correlated to their electronic configurations. Only three such families viz., noble gases, the halogens and alkali metals have been taken up for study in this chapter.

#### Outline

1. Noble gases (7.1)
2. Halogens (7.2)
3. Alkali metals (7.3)
4. Electronic configuration and similarity of properties (7.4)

#### Concepts

The properties of elements depend on their electronic configurations.

#### SCHEDULE

Lesson	Experiment	Postlab	Demonstration	Text covered
1	—	—	I	7.1
2	—	Discussion	II	7.2
3	—	Discussion	II	7.2
4	—	—	—	7.2
5	16	Discussion	—	7.2
6	—	—	—	—
7	17	Discussion	III	7.2
8	18	Discussion	—	7.2
9	—	Discussion	IV	7.3
10	19	Discussion	V	7.3
11	20, 21	Discussion	—	7.4
12	—	—	—	Review & quiz

## Development

### 7.1 Noble gases

Recall that air contains besides oxygen, nitrogen, carbon dioxide and traces of water vapour, minute quantities of neon, argon, krypton, xenon and helium, called noble gases. Exhibit the tubes filled with these gases. Let the students note that these gases are colourless. They produce characteristic colours when an electric discharge is passed through them.

#### Demonstration I : Discharge of electricity through noble gases

Time required : 10 minutes

#### Materials required

Accumulator or battery (12 V)

Induction coil

Sealed discharge

Rheostat

tubes of helium

Switch

and neon

Connecting wires

#### Procedure

Set up the arrangement as shown in Fig 7.1. Connect the discharge tube containing helium gas to the secondary terminals of the induction coil. Put the switch on and let the students observe and record the characteristic pink colour produced by the helium gas. Repeat the procedure with the tube containing neon gas. Let the students observe and record the crimson red colour produced by the neon gas.

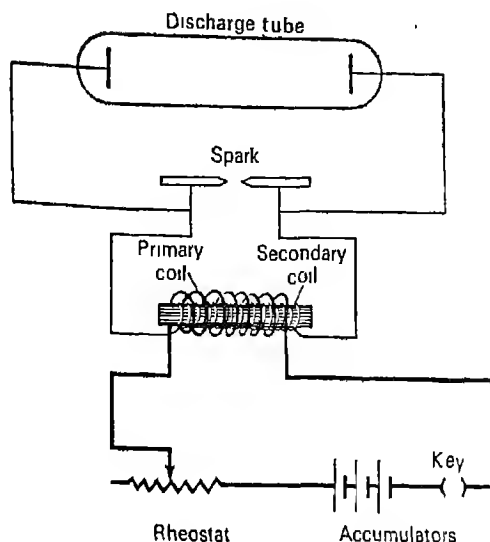
Point out to the students the use of these gases particularly the neon signs for advertisement purposes. In actual practice mixtures of these gases are generally used to give different shades of colour.

Recall the burning of magnesium in air and point out that the products obtained are  $\text{MgO}$  with a little  $\text{Mg}_3\text{N}_2$ . The noble gases present in the air do not react with magnesium even at high temperature.

It is observed that air free from carbon dioxide and moisture, when subjected to electric sparks (2500–3000°C) yields nitric oxide. The residual gas left after the reaction, mainly contains noble gases. These two experiments indicate that the noble



gases are chemically inert. However, compounds of xenon and krypton have been prepared recently.



*Fig. 7.1 Discharge of electricity through noble gases*

## 7.2 Halogens

Discussion on halogens can begin with Demonstration II.

### **Demonstration II : Reactions of halogens**

Time required :  $1\frac{1}{2}$  hours

#### **Materials required**

Gas collecting jars with glass plates	18	Jars filled with chlorine, bromine, iodine, hydrogen and hydrogen sulphide
Glass rod		
Deflagrating spoons	2	Sodium
Water bath		Potassium permanganate Concentrated hydrochloric acid Zinc pieces Iron (II) sulphide Dilute hydrochloric acid Aqueous ammonia

Bromine

Iodine

Litmus papers

Starch paper

- (i) Before starting the demonstration, fill the gas jars (4 each) with chlorine, bromine (about 2 drops in each) and iodine (about 0.2 g in each). It is better to warm the gas jars containing iodine by keeping them on a hot water bath so that the violet vapours of iodine can be clearly seen.
- (ii) Keep three jars filled with hydrogen.
- (iii) Keep three gas jars filled with hydrogen sulphide. Prepare hydrogen sulphide in a Kipp's apparatus using iron (II) sulphide and dilute hydrochloric acid. Collect the gas in gas jars by upward displacement of air.

### Procedure

- (1) Exhibit to the students the jars containing halogens and ask them to record the colour of each.

Chlorine      -      Greenish yellow

Bromine      -      Reddish brown

Iodine        -      Violet

- (2) Introduce moistened litmus papers (red and blue) into one of the jars of chlorine, bromine and iodine. Let the students note the change in the colour of litmus. Elicit from the students that chlorine bleaches litmus readily. In bromine the bleaching is slow and in iodine least.
- (3) Repeat the above test with starch paper. Let the students conclude that starch paper turns blue only in the jar containing iodine vapour.
- (4) Place a small piece of sodium on the deflagrating spoon and plunge it into a jar of chlorine. Let the students note the vigour of the reaction. Let them also observe that the colour of chlorine will disappear with the formation of sodium chloride.

Repeat the above test with bromine and iodine. Let the students note that the reaction is not appreciable in cold. Now warm the sodium metal and introduce it into the jars of bromine and iodine.

In bromine the reaction is slow while in iodine it is much more so. Elicit from the students that reactivity decreases from chlorine to bromine to iodine.

**Precaution :** Any unreacted sodium left on the deflagrating spoon should be destroyed by plunging it in a trough of cold water (care !).

- (5) Invert the jars of chlorine, bromine and iodine simultaneously over three jars of hydrogen (caution! in the case of chlorine avoid direct exposure to sunlight since there is a danger of explosion). Let the students observe the disappearance of the colour of the halogens in each case. Colour disappears first in chlorine, then in bromine and then in iodine. Test the products formed in each case by introducing a glass rod dipped in ammonia solution. White fumes are observed more in chlorine, less in bromine and no fumes are observed in iodine.
- (6) Invert three jars of hydrogen sulphide over three jars of chlorine, bromine and iodine. Let the students note the disappearance of the colour and the deposition of the pale yellow precipitate of sulphur on the sides of the gas jars. Elicit from the students that more sulphur is deposited in the case of chlorine, less in bromine and least in iodine.

Now the students can perform experiment 16.

### **Experiment 16 : Some reactions of aqueous solutions of chlorine, bromine and iodine**

Time required : 1 hour

The solutions required for this experiment are prepared as follows :

Chlorine solution : Pass chlorine gas through distilled water till the solution smells distinctly of chlorine.

Bromine solution : Dissolve 1 ml of bromine in 200 ml of distilled water.

Iodine solution : Dissolve about 0.25 g of solid iodine in 50 ml of ethyl alcohol and then dilute to 350 ml with distilled water.

Iron (II) sulphate solution : Iron (II) sulphate is easily oxidized in air. Iron (II) ammonium sulphate (Mohr's salt) is comparatively more stable. So prepare a solution of Mohr's salt by dissolving 4 g of the salt in 1 litre of distilled water using 10 ml dilute sulphuric acid.

Ammonium thiocyanate solution. Dissolve 1 g of ammonium thiocyanate,  $\text{NH}_4\text{CNS}$ , in 100 ml of distilled water.

**Hydrogen sulphide solution** : Prepare an aqueous solution of hydrogen sulphide by bubbling the gas through distilled water until the solution smells distinctly of the gas. However, this solution does not keep for long.

**Starch paper** : Boil about 100 ml of water and add to it slowly with stirring a paste of 1 g of starch (soluble) with a few ml of water. After addition continue heating for about 5 minutes and then allow it to cool. Dip filter papers in this solution and dry in air.

Results are given in the following table :

Reaction	Chlorine solution	Bromine solution	Iodine solution
1. Colour	Greenish yellow	Reddish brown	Brown
Smell	Irritating	Irritating	Irritating
Litmus papers and petals of flowers	Bleaching action is quick	Bleaching action is feeble	Almost Nil
Starch paper	No action	Turns yellow	Turns blue
2. Carbon tetrachloride	Very pale greenish yellow	Orange red	Violet
3. Iron (II) ammonium sulphate and ammonium thiocyanate	Deep blood red colour	Red colour	Very light pink colour
4. Hydrogen sulphide solution	White precipitate of sulphur on standing	White precipitate of sulphur on standing	—

### Postlab discussion

Elicit from the students the following :

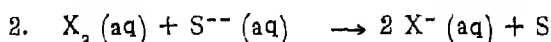
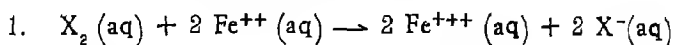
- Iron (III) sulphate gives red colour with ammonium thiocyanate solution but iron (II) sulphate does not; but on addition of chlorine solution a deep blood red colour is produced. This is due to the oxidation of iron (II) to iron (III) by chlorine.

- (ii) On repeating the experiment with bromine solution, a red colour is produced but with less intensity.
- (iii) On repeating the experiment with iodine solution, very light pink colour is produced.

chlorine > bromine > iodine.

### Questions and answers :

1. Write ionic equations for the reactions of chlorine, bromine and iodine with iron (II) sulphate and hydrogen sulphide.



where  $X = Cl$  or  $Br$

2. Name a colour test given by iodine but not by chlorine.

Starch solution turns blue when treated with iodine.

3. Carbon tetrachloride, when shaken with bromine and iodine solutions, becomes coloured. Why?

Bromine and iodine dissolve readily in carbon tetrachloride.

Now Demonstration III may be taken up to show the reactivity of halogens.

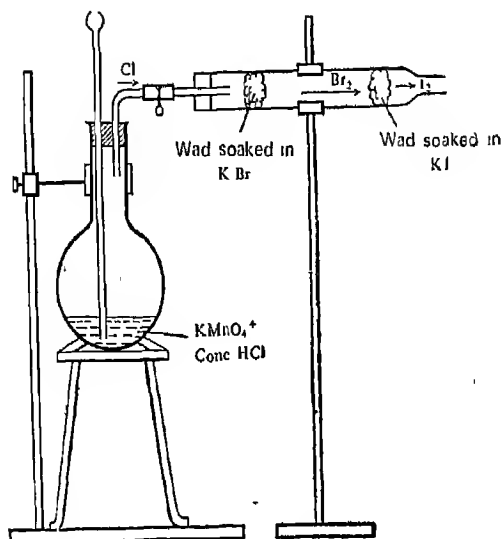
### Demonstration III : Reactivity of halogens

Time required : 10 minutes

#### Materials required

Chlorine generator	Potassium permanganate
Glass tube ( 1 X 2 cm)	Concentrated hydrochloric acid
Cotton wool :	Potassium bromide solution (5%)
	Potassium iodide solution (5%)

Before starting the demonstration keep the assembly of apparatus as shown in Fig. 7.2 ready.



*Fig. 7.2 Reactivity of halogens*

### Procedure

Place a wad of cotton wool soaked in potassium bromide solution at one end of the glass tube and another soaked in potassium iodide at the other end. Allow chlorine gas to pass into the tube. Elicit from students how brown fumes of bromine are liberated from the first wad of cotton wool. Let them also observe that the second turns brown due to the liberation of iodine. Point out that chlorine displaces bromine from potassium bromide and bromine displaces iodine from iodide. Arrange the three elements in order of reactivity.

Now the students can proceed to perform experiments 17 and 18.

### Experiment 17 : Some reactions of chloride, bromide and iodide ions.

Time required :  $1\frac{1}{2}$  hours

**Silver nitrate solution :** Prepare 2% silver nitrate solution using distilled water.

**Starch iodide paper :** Prepare by dipping filter papers first in potassium iodide solution (5%) and then in starch solution.

**Lead nitrate solution :** Prepare a 5% lead nitrate solution in distilled water.

The results are given in the following table :

Reagent	NaCl	KBr	KI
1. Conc. $\text{H}_2\text{SO}_4$	Colourless gas (HCl)	Colourless gas (HBr) followed by reddish brown fumes of bromine	Colourless gas (HI) followed by violet fumes of iodine
Smell	Pungent	Choking	Choking
Glass rod moistened with ammonia	Dense white fumes	White fumes	...
Starch paper	No change	Turns yellow	...
Starch iodide paper	...	Blue	...
Moist blue litmus paper	Turns red	Turns red	Turns red
2. $\text{MnO}_2$ and conc. $\text{H}_2\text{SO}_4$	Greenish yellow gas (chlorine)	Reddish brown fumes (bromine)	Violet fumes (iodine)
Glass rod dipped in aq. ammonia	White fumes	...	...
Starch paper	...	Turns yellow	Turns blue
Starch iodide paper	Turns blue	Turns blue	Turns blue
Moist litmus papers	Decolourised (both red and blue)	Decolourised slowly	No change
3. Silver nitrate	Curdy white precipitate	Pale yellow precipitate	Yellow precipitate
Addition of dil. $\text{HNO}_3$ to the precipitate	Insoluble	Insoluble	Insoluble
Addition of aq. ammonia	Soluble	Soluble with great difficulty	Insoluble
4. Lead nitrate	White precipitate	White precipitate	Yellow precipitate
Precipitate + boiling water	Soluble; reprecipitated on cooling	Soluble; reprecipitated on cooling	Soluble; on cooling golden yellow spangles

### **Postlab discussion**

The students should be advised not to smell chlorine, bromine and iodine vapours.

Sodium chloride when heated with concentrated sulphuric acid in the presence of manganese dioxide yields chlorine gas. However, some hydrogen chloride gas may remain unoxidized. This causes some white fumes when a glass rod dipped in aq. ammonia is brought near the gas.

While testing the solubility of lead halides in water, the students may be asked to take a small amount of the precipitate and proceed.

Point out to the students that precipitation of halides with silver nitrate solution will serve as a test for the detection of halide ions.

Emphasise that the halides of lead are soluble in hot water and are reprecipitated in crystalline form on cooling.

### **Experiment 18 : Reactivity of chlorine, bromine and iodine**

**Time required :** 40 minutes.

5 per cent solutions of KCl, KBr and KI to be used.

Aqueous solutions of halogens to be prepared as in experiment 16.



Reagent	KCl solution	KBr solution	KI solution
1. Chlorine solution	No colour change	Brownish yellow due to the liberation of bromine	Reddish brown due to the liberation of iodine
Addition of carbon tetrachloride to above	No change (light greenish yellow due to chlorine solution added)	Red globule	Dark violet globule
2. Bromine solution	No change (yellow colour due to bromine solution added)	No change (yellow colour due to bromine solution added)	Light brown due to the liberation of iodine
Addition of carbon tetrachloride to above	No change (red globule due to bromine solution added)	No change (red globule due to bromine solution added)	Violet globule due to iodine
3. Iodine solution	No change (brown due to iodine solution added)	No change (brown due to iodine solution added)	No change (brown due to iodine solution added)
Addition of carbon tetrachloride to above	No change (violet due to iodine solution added)	No change (violet due to iodine solution added)	No change (violet due to iodine solution added)

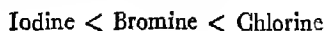
### Postlab discussion

It may be mentioned that the solubility of halogens in carbon tetrachloride is more than that in water. Therefore, when carbon tetrachloride is added to aqueous solutions of halogens, the halogen is preferentially transferred to the carbon tetrachloride layer imparting characteristic colour.

Point out to the students that chlorine displaces bromine from potassium bromide and iodine from potassium iodide. Bromine displaces only iodine from potassium iodide. Iodine does not displace either chlorine or bromine. Thus chlorine is the most reactive and iodine is the least.

### Answer to Question

Based on your observations, arrange the elements chlorine, bromine and iodine in the increasing order of their reactivity.



### 7.3 The Alkali metals

Now the properties of alkali metals may be taken up through Demonstration IV.

#### Demonstration IV : Some properties of alkali metals

Time required : 30 minutes

#### Materials required

Pen knife	Bottle containing sodium and
Watch glass	potassium under kerosene
Tongs	Litmus papers
Gas jars filled with chlorine-2	Filter papers
Trough	
Deflagrating spoon	

Before starting the demonstration, have two jars filled with chlorine as in demonstration II.

#### Procedure

- (i) Take out a piece of sodium, dry to free of kerosene between the folds of a filter paper, keep it on the watch glass and cut it with the pen knife. Let the students observe (i) the ease with which sodium can be cut and (ii) the brightness of the freshly cut surface. Let them note the tarnishing of the surface on keeping it exposed to air.
- (ii) Warm a small piece (pea-size) of sodium on a deflagrating spoon and introduce it into the jar of chlorine. Call the attention of the students to the white deposit formed on the walls of the jar.
- (iii) Throw a small piece (CAUTION) into a trough of water. Call the attention of the students to (i) the fact that the metal floats on water (indicating that its density is lower than that of water) and (ii) to the colour of the flame (blue) due to the burning of hydrogen liberated in the reaction.

After the reaction is over, test the solution in the trough with litmus paper. Elicit from the students that the solution is alkaline as it turns red litmus paper blue.

- (iv) Repeat the above tests with potassium.

Bring out clearly the similarity in the behaviour of sodium and potassium. If conducted under more controlled conditions (e.g. the area of surface of water, the surface of the metal piece and the temperature remaining the same), it is possible to bring out that potassium is more reactive than sodium.

The students can now perform experiment 19.

### Experiment 19 : Some reactions of alkali metal hydroxides

Time required : 30 minutes

Aqueous solutions of the following concentration may be used.

1. Sodium hydroxide ... .. 5%
2. Potassium hydroxide ... .. 5%
3. Copper sulphate ... .. 2.5%
4. Zinc sulphate ... .. 2.5%
5. Magnesium sulphate .. ... 1.5%
6. Ammonium chloride ... .. 2.5%
7. Iron (III) chloride .. ... 1%

The results are given in the following table.

Reagent	Copper (II) sulphate solution	Iron (III) chloride solution	Zinc sulphate solution	Magnesium sulphate solution
NaOH	Blue precipitate	Reddish brown precipitate	White precipitate	White precipitate
NaOH in excess	Precipitate insoluble	Precipitate insoluble	Precipitate insoluble	Precipitate insoluble
KOH	Blue precipitate	Reddish brown precipitate	White precipitate	White precipitate
KOH in excess	Precipitate insoluble, on standing changes to green and finally black	Precipitate insoluble	Precipitate is soluble	Precipitate insoluble

Both sodium hydroxide and potassium hydroxide on exposure to air absorb moisture and dissolve to form a solution. When ammonium chloride is heated with sodium hydroxide, ammonia gas with its characteristic smell is evolved. It turns red litmus paper blue, gives dense white fumes with a glass rod dipped in concentrated hydrochloric acid. Point out to the students that sodium hydroxide and potassium hydroxide liberate ammonia from any ammonium salt.

## Answer to Question

Explain what happens to sodium hydroxide when it is exposed to air.

On exposure to air, sodium hydroxide absorbs moisture and dissolves in it to form a solution. Thus it is deliquescent.

Before the students take up experiments 20 and 21 demonstrate the correct method of carrying out the flame test.

## Demonstration V : Flame colourations of sodium and potassium salts

Time required : 10 minutes

### Materials required

Platinum wire (5 cm long)	Sodium carbonate 0.5 g
Watch glass	Potassium carbonate 0.5 g
Glass rod	Concentrated hydrochloric acid

A thin platinum wire about 5 cm. long, fused into the end of a short piece of glass tubing or glass rod which serves as a handle, is employed.

### Procedure

First thoroughly clean the platinum wire by dipping it into concentrated hydrochloric acid, contained in a watch glass and then heating it in the non-burning part of the flame till no colour is imparted to the flame.

Place a small quantity of the substance i. e. sodium carbonate on a watch glass, moisten with a little concentrated hydrochloric acid, (explain to the students that the salts are first converted to the chlorides, since the chlorides are the most volatile of the salts) and take a little of the substance on a clean platinum wire and introduce into the base of the nonluminous Bunsen flame. Observe the colour imparted to the flame.

Clean the platinum wire and repeat the test with potassium carbonate.

Metal	Colour
Sodium	Golden yellow flame
Potassium	Lilac flame

Recall to the students the colours of the sparklers which they might have lit up during Diwali. Mention to them that these colours are due to mineral salts barium (apple green), copper (blue), strontium (crimson), calcium (brick red), etc.

The students now can perform experiments 20 and 21.

**Experiment 20 : Some reactions of alkali metal salts**

Time required : 20 minutes.

The results are given in the following table.

	Substances					
	Sodium carbonate (washing soda)	Sodium sulphate (anhydrous)	Sodium nitrate	Potassium carbonate	Potassium sulphate	Potassium nitrate
<b>Reactions</b>						
1. Testing vapours with cobalt chloride paper	Turns pink	No change	No change	Turns pink	No change	No change
2. Testing with glowing splinter	—	—	Rekindled	—	—	Rekindled
3. Aqueous solution tested with litmus paper	Turns blue	—	—	Turns blue	—	—

**Experiment 21 : Flame colourations of alkali metals.**

Time required : 15 minutes.

The results are given in the following table.

Metal salts	Flame colouration	Flame colouration through cobalt glass
Sodium salts	Golden yellow	Nil
Potassium salts	Lilac	Crimson

In the procedure given in the laboratory manual, the students are directed to make a paste of sodium chloride in water. It should be pointed out to the students

that in the case of substances other than metal chlorides, the substance is to be treated with a few drops of concentrated hydrochloric acid and then the flame test carried out

### **Postlab discussion.**

Point out to the students that the flame colour is due to the metal ions (cations) and not anions. Hence all sodium salts give the same characteristic golden yellow flame. All potassium salts give lilac flame.

### **7.4 Electronic configuration and properties.**

The students may now be taught the text material under 7.4 to give the explanation for the similarity of properties in the group of elements so far described.

## BACKGROUND INFORMATION

### 7.1 Noble gases.

The six elements that are not chemically active to any significant degree are called the rare or noble gas elements. These gases are helium, neon, argon, krypton, xenon and radon. Except argon, which is third in abundance among all gases in the atmosphere, they exist only in minute amounts (Table 7.1). Radon is a radioactive element, being created during the decay of radium and in turn decays into Radium-A (Po) with a half life of about 3.8 days.

Table 7.1 Composition of Dry Air.

Substance	Per cent by volume
N <sub>2</sub>	78
O <sub>2</sub>	21
Ar	0.93
CO <sub>2</sub>	Varies - average 0.03
H <sub>2</sub>	0.01
Ne	0.0018
He	0.0005
Kr	0.0001
Xe	0.0001

Each element, except helium, has eight electrons in the highest filled energy shell; two electrons are in the s subshell, and six are in the p subshell. Helium has a complete first energy shell with two electrons in the 1s shell. There is no tendency for these elements to combine with one another, and they all exist as monatomic gases.

### The discovery of the noble gases.

Argon was the first of the noble gases to be isolated. The story begins with the experiments of Henry Cavendish; in 1785. He passed electric sparks through a mixture of oxygen and air in the presence of an alkaline solution. All of the nitrogen in the mixture was converted into the oxides of nitrogen and absorbed by the alkali, but a part of the air failed to be oxidized and absorbed. He estimated that after the removal of the excess of oxygen by alkali sulphides there still remained a

small residue of gas, amounting to  $1/120$  of the original volume of the air. The residual gas differed from oxygen and nitrogen. In 1894 Lord Rayleigh during a systematic investigation on the density of gases found that the density of nitrogen, prepared from the atmospheric air, was slightly higher (about 0.5 per cent) than the density of nitrogen prepared chemically from its compounds. This led to the conclusion that the atmospheric nitrogen is not pure and probably contains some other constituents which are heavier than nitrogen. Sir William Ramsay continued the investigation. He passed the nitrogen obtained from air over red hot magnesium to find out whether or not it would be completely absorbed by combination with the metal to form magnesium nitride. He found a residual gas,  $1/80$  of the volume of nitrogen used, which would not combine with any other element. The gas was about  $15/14$  as heavy as the original nitrogen. When Ramsay examined the spectrum of this gas, he saw not only the bands of nitrogen but also groups of red and green lines never before observed in the spectrum of any gas. In August 1894, Ramsay and Rayleigh announced the discovery of the first inert gas, which they called argon, the lazy one.

Helium, the lightest of the rare gases, was actually identified some years before the discovery of argon. It was first found in the sun. Its presence there was discovered by means of spectroscopy. In 1868, the French astronomer Pierre Janssen came to India to observe a total eclipse of the sun and to make a spectroscopic study of its chromosphere. He noticed a yellow line that did not coincide with the D-line of sodium. Later that year Sir Joseph Lockyer, the English astronomer, using a specially designed spectroscopy, verified Janssen's discovery. He found that this new yellow line did not belong to any element then known. Therefore he attributed this line to the presence of a new element in the sun, and named it helium from the Greek meaning "sun". Helium occurs in certain radioactive minerals.

Ramsay and co-workers isolated three other rare gases from liquid air samples between 1894 and 1898. These were named neon, from the Greek word for new; Krypton, from the Greek word for hidden; and xenon, from the Greek word for strange. The last of the rare gases to be discovered was radon. It was isolated as one of the radioactive disintegration products of radium, from which it was named.

### Isolation

The noble gases are isolated from liquid air by fractional distillation.

The boiling points of the various constituents of liquid air are given in table 7.2.

Table 7.2: The boiling points of the various constituents of liquid air

Xenon	...	...	$-107.1^{\circ}\text{C}$
Krypton	...	...	$-152.9^{\circ}\text{C}$
Oxygen	...	...	$-182.9^{\circ}\text{C}$
Argon	..	..	$-185.85^{\circ}\text{C}$
Neon	...	...	$-245.92^{\circ}\text{C}$
Nitrogen	...	...	$-195.7^{\circ}\text{C}$
Helium	...	...	$-268.87^{\circ}\text{C}$



## Uses.

Helium is preferred to hydrogen for filling balloons since there is no risk of explosion. Helium mixed with oxygen is used as artificial air in place of air for deep-sea divers. Helium is less soluble in the blood than is nitrogen, and is therefore more rapidly eliminated when the diver comes to the surface. Thus helium containing atmosphere prevents the formation of gas bubbles in the blood (the bends), which develop on rapid decompression. In metallurgy, particularly in welding metals like aluminium, magnesium or its alloys which have a tendency of getting easily oxidized helium and argon are largely used for the production of an inert atmosphere.

Argon is used for filling tungsten electric-light bulbs, for it produces greater brilliance and prolongs the life of the bulbs. The familiar crimson glow of neon lighting is made by passing an electric discharge through a glass tube containing a small quantity of neon and mercury vapour at low pressure. Mercury vapour, neon, helium, and argon are used in this type of lighting both separately and in various combinations. Tubes containing krypton and xenon have been developed for use in high speed photography.

Because of its radioactive property radon is largely employed in radio-therapy for treating malignant growths. It is used in the treatment of cancer.

## Compounds of noble gases.

For many years attempts were made to combine the noble gases with other elements. The results were inconclusive. Finally in 1962, the element xenon was found to form a compound, xenon hexafluoroplatinate,  $\text{XePtF}_6$ . Since then, compounds of xenon and krypton have been prepared, e.g., xenon difluoride,  $\text{XeF}_2$ , xenon tetrafluoride,  $\text{XeF}_4$ , xenon hexafluoride,  $\text{XeF}_6$ , krypton difluoride  $\text{KrF}_2$ , etc. (Table 7.3).

Table 7.3. Some compounds of the noble gases.

Compound	$\text{KrF}_2$	$\text{XeF}_2$	$\text{XeF}_4$	$\text{XeF}_6$
Physical form	Colourless crystals	Same	Same	Same
Melting point	Sublimes below $0^\circ\text{C}$	$140^\circ$	$114^\circ\text{C}$	$48^\circ\text{C}$
Reactivity	Decomposes spontaneously at room temperature	Reacts with water to form $\text{Xe}$ and $\text{O}_2$	Stable	Stable

## 7.2 Sources of halogens.

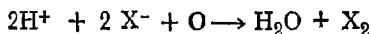
Fluorine does not occur free in nature because it is highly reactive. The important fluorine containing minerals are fluorspar,  $\text{CaF}_2$ , and cryolite,  $\text{Na}_3\text{AlF}_6$ . Fluorspar is the chief source of fluorine compounds.

Chlorine, also being a reactive element, occurs in nature only in the combined state. It is found in combination with sodium, magnesium and several other metals. The most common compound is sodium chloride. Sea water contains about 2.5% by weight of sodium chloride, together with small amounts of other dissolved salts.

Bromine is far less abundant than chlorine. It occurs in sea water in the form of the soluble bromides,  $\text{NaBr}$ ,  $\text{KBr}$  and  $\text{MgBr}_2$  in small amounts (less than 0.001 per cent). Certain marine plants and animals absorb it from sea water. So it is found in them in greater concentration, e.g., in sea weeds and oysters. Iodine also occurs as sodium iodate ( $\text{NaIO}_3$ ) in Chile salt petre ( $\text{NaNO}_3$ ), forming about 0.2 per cent of the total salt. Iodine also occurs in sea weeds.

### Preparation of the halogens.

All the halogens except fluorine are prepared by treating the halide ion with an appropriate oxidizing agent. The general reaction may be expressed by the following schematic oxidation - reduction equation, where  $\text{X}^-$  represents the halide ion and  $\text{X}_2$  the halogen.



Among the oxidizing agents used are manganese dioxide, potassium permanganate, potassium chlorate, and nitric acid. Of these manganese dioxide is most commonly employed.

Fluorine cannot be obtained directly by chemical methods due to its high reactivity. It is prepared by the electrolysis of a mixture of fused potassium hydrogen fluoride and anhydrous hydrogen fluoride. The electrolysis is carried out in a copper vessel with graphite electrodes. Hydrogen fluoride is non-conducting but when dissolved in fused potassium fluoride, it gives a conducting solution. Fluorine reacts with copper and forms a protective layer of copper fluoride on the surface. Fluorine is liberated at the anode and hydrogen at the cathode.

## Uses of halogens

Fluorine is used in the preparation of several useful compounds. Freon,  $\text{CCl}_2\text{F}_2$ , is used as a refrigerant because of its nonpoisonous and non-corrosive nature. Teflon, a polymer of tetrafluoroethylene,  $\text{C}_2\text{F}_4$ , is inert to chemical attack and so finds many applications. Hydrogen fluoride is used for etching glass. Fluorine in trace amounts is added to drinking water in some countries, as it helps protect dental enamel from premature decay.

Chlorine is used in the manufacture of organic compounds such as chloroform and carbon tetrachloride. It is also used in the manufacture of plastics and synthetic rubber. Chlorine is used for sterilizing water. It is used as a bleaching agent.

Large quantities of bromine are used to produce ethylene bromide,  $\text{C}_2\text{H}_4\text{Br}_2$ , which is added to petrol along with lead tetraethyl,  $\text{Pb}(\text{C}_2\text{H}_5)_4$ , as an 'anti-knock' compound. During combustion, lead tetraethyl decomposes and forms lead deposits in the engine. The use of ethylene bromide prevents the formation of lead deposits. The bromine is also used in the preparation of silver bromide (a light-sensitive material) which is used in the preparation of photographic plates and films.

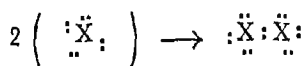
The most important use of iodine is as an antiseptic. A solution of iodine in ethyl alcohol, known as "tincture of iodine" is marketed for domestic use. Iodine is also used in the preparation of a number of metal iodides and organic compounds. Silver iodide is used in photography.

## General properties of the halogens

The electronic configurations of these elements along with other properties are listed in Table 7.4. In all these elements the outermost shell contains seven electrons ( $s^2 p^5$ ). Each of the halogens differs from the noble gas that follows it in the periodic system by having one electron less in its valence shell.

Because of their high electronegativity, the halogens show practically no metallic properties, though solid iodine has a somewhat metallic appearance. Astatine, the heaviest member of the group may also have some metallic properties, but it is a short-lived radioactive element. Halogens attain the stable configuration of the noble gas type either by gaining one electron to form a univalent negative ion or by sharing a

pair of electrons to form a single covalent bond. The former occurs when a halogen atom combines with an atom of low electronegativity to form an ionic bond as in NaCl or KCl. The latter occurs when the halogen atom combines with another atom of comparable electronegativity to form a covalent bond as in Cl<sub>2</sub>. The diatomic halogen molecules can be represented in the following manner :



X = halogen

The melting or boiling point of the halogens increases regularly with the increase in their atomic weights

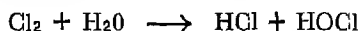
The halogens have a tendency to gain electrons and can therefore act as oxidizing agents. The oxidizing action decreases with decreasing electronegativity, i.e. from fluorine to astatine. Fluorine is the strongest oxidizing agent and iodine (and more so astatine) is the weakest oxidizing agent.

Table 7.4 Some properties of the halogens.

	F Fluorine	Cl Chlorine	Br Bromine	I Iodine
Colour	Light yellow gas	Greenish yellow gas	Reddish brown liquid	Dark violet lustrous solid
Atomic number	9	17	35	53
Electronic structure	2, 7	2, 8, 7	2, 8, 18, 7	2, 8, 18, 18, 7
Atomic weight	19.00	35.46	79.92	126.9
Melting point (°C)	- 222	- 102	- 7.3	114
Boiling point (°C)	- 187	- 34.6	58.8	183
Electronegativity (Pauling scale)	4.0	3.0	2.8	2.4

## Solubility of halogens in water and other solvents

Fluorine reacts with water instantaneously giving fluorine monoxide,  $F_2O$ , and hydrogen fluoride (some hydrogen peroxide, ozone and oxygen are also formed). Chlorine is fairly soluble in water. Dissolved chlorine reacts slowly with water to give hydrochloric (HCl) and hypochlorous (HOCl) acids.



An aqueous solution of chlorine, often called chlorine water, is an effective bleaching and sterilizing agent because both the hypochlorous acid and dissolved chlorine are oxidizing agents.

Chlorine is also soluble in carbon tetrachloride, chloroform and carbon disulphide.

Bromine is moderately soluble in water but very soluble in carbon tetrachloride, carbon disulphide and chloroform forming orange red or brown solutions.

Iodine is very sparingly soluble in water but is soluble in aqueous solution of potassium iodide giving a dark brown solution. It dissolves readily in carbon disulphide, chloroform and carbon tetrachloride to give a violet or purple solution.

Thus, the solubility in water of these three elements decreases from chlorine to iodine.

## Halides in photography

Silver chloride and silver bromide are used in photographic industry. Both of these compounds are light sensitive, being slowly decomposed into the elements by light. Silver bromide is crushed to a very fine powder and suspended in gelatin to form an emulsion, which is then coated on a thin sheet of celluloid or other plastic to make a photographic film. When this film is exposed to light, the crystal lattice of the AgBr is disrupted and made sensitive to chemical attack in proportion to the intensity and time extent of light exposure. If the light is focused on the film through a lens as in a camera, a latent image is formed in the film.

The film is then *developed* by immersing it in a solution of a weak reducing agent, known as the photographic *developer*, such as hydroquinone. The developer reduces the crystals of silver halide which have been attacked by light (during the formation of the latent image) to metallic silver and so forms a visible image. The unexposed silver halide is not affected by the developer and is later removed by immersing the film in a *fixing bath* of sodium thiosulphate. At this stage, the film contains metallic silver in

those areas previously affected by light in the original exposure. The film is now darkest in those parts where it was most intensely affected by light, and is hence called a *negative*. A positive print is prepared by shining light through a developed negative upon a piece of photosensitive paper (coated with an emulsion of silver bromide). The print is then developed just as the film was and dried.

### 7.3 The alkali metals

#### Occurrence and preparation

Alkali metals being very reactive do not occur free in nature. They are always found in the form of their compounds. Compounds of sodium and potassium occur very abundantly in nature whereas those of other alkali metals, rubidium and cesium occur only sparingly. All the alkali metals occur as compounds of mixed oxides or halides. For example, potassium occurs abundantly as the mineral feldspar,  $K_2O$ ,  $Al_2O_3$ ,  $6SiO_2$ . Lithium occurs in nature in several types of rocks of very complex composition. The most abundant compound of sodium in nature is sodium chloride.

Alkali metals are generally prepared by the electrolysis of molten alkali metal chlorides. For example, metallic sodium is prepared industrially by the electrolysis of a fused mixture of  $NaCl$  and  $CaCl_2$  at about  $600^\circ C$ . Calcium chloride is added to lower the melting point of the sodium chloride. Sodium metal is formed at the iron or copper cathode, and chlorine at the carbon anode.

#### General properties of the alkali metals

The electronic configuration, of alkali metals along with their physical properties are listed in Table 7.5. It may be noted from the table that in each element the neutral atom has one electron more than the preceding noble gas.

Table 7.5. Some physical properties of alkali metals

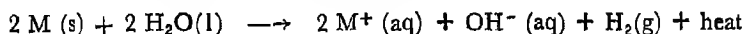
	Li	Na	K	Rb	Cs
Atomic number	3	11	19	37	55
Electronic configuration	2, 1	2, 8, 1	2, 8, 8, 1	2, 8, 18, 8, 1	2, 8, 18, 18, 8, 1
Melting point ( $^\circ C$ )	186	97.5	62.3	38.5	28.5
Boiling point ( $^\circ C$ )	1336	880	760	700	670
Electronegativity (Pauling scale)	1.0	0.9	0.8	0.8	0.7
Flame colour	Red	Yellow	Lilac	Red	Blue

The alkali metals have a bright lustre when freshly cut. They are excellent conductors of heat and electricity. They are malleable.

Alkali metals are soft and have low melting points. The softness increases and the melting and boiling points decrease in the group from lithium to cesium. Thus lithium is the hardest member of the group and has the highest melting and boiling points whereas cesium is very soft and has the lowest melting and boiling points.

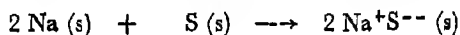
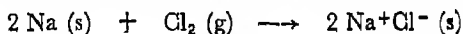
The alkali metals have a great tendency to give up their valence electron to electronegative elements like oxygen, sulphur, halogens etc., to form stable ionic compounds.

Another important reaction of the alkali metals is that they react with water to release hydrogen and form basic and highly water soluble hydroxides (except for lithium hydroxide which is only sparingly soluble).



The rate of reaction is slowest with lithium and fastest with cesium. When rubidium or cesium reacts with water enough heat is produced to cause the liberated hydrogen gas to ignite. These metals, therefore, must be stored under some inert liquid such as kerosene.

As alkali metals readily lose their single valence electron, they act as strong reducing agents. Any oxidizing agent can be reduced by the alkali metals.



The alkali metals as their compounds impart characteristic colours to a gas flame when heated in it. For example, lithium gives red, sodium yellow, potassium lilac, rubidium red and cesium blue coloured flames. Solution of chlorides are generally used in performing flame tests, since they are more volatile than most other compounds of these elements. The colours of the flame are utilised in the detection of these elements.

Each question is followed by four answers. One of them is correct. Select the correct answer and indicate it in the bracket provided.

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7.  $\text{MnO}_2(\text{s}) + 4 \text{H}^+(\text{aq}) + \text{XCl}^-(\text{aq}) \longrightarrow \text{Mn}^{++}(\text{aq}) + \text{Cl}_2(\text{g}) + \text{YH}_2\text{O}(\text{l})$ .
- In the above equation the values of X and Y are respectively ( )
- (a) 2, 1 (b) 1, 2  
(c) 4, 1 (d) 2, 2.
8. In the preparation of chlorine, manganese dioxide,  $\text{MnO}_2$ , functions as ( )
- (a) a catalyst (b) an oxidizing agent  
(c) a reducing agent (d) a drying agent.
9. An HCl generator contains NaCl and  $\text{H}_2\text{SO}_4$ . Which of the following reagents should be added to convert it in to a chlorine generator? ( )
- (a) KCl (b)  $\text{Br}_2$   
(c)  $\text{MnCl}_2$  (d)  $\text{MnO}_2$ .
10. Preparation of chlorine gas from NaCl involves ( )
- (a) neutralization (b) oxidation  
(c) reduction (d) precipitation.
11. With what elements among the following does chlorine react most readily? ( )
- (a) carbon (b) hydrogen  
(c) nitrogen (d) oxygen.
12. Of the following the strongest oxidizing agent is ( )
- (a) chlorine (b) chloride ion  
(c) bromine (d) iodine.
13. Of the following the strongest reducing agent is ( )
- (a) chlorine (b) chloride ion  
(c) bromine (d) iodine.

14. The increasing order of reactivity of alkali metals is ( )  
(a) K, Li, Na (b) Li, K, Na  
(c) Li, Na, K (d) Na, K, Li.
15. The atom which has the electronic configuration 2, 8, 1 is that of ( )  
(a) neon (b) sodium  
(c) chlorine (d) potassium.
16. Five elements have atomic numbers 2, 9, 10, 11, 18 respectively. Of these the elements with the following atomic numbers will have similar properties. ( )  
(a) 2, 9, 18 (b) 9, 11, 18  
(c) 2, 10, 18 (d) 2, 9, 11.

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## Periodic Classification of Elements

### INSTRUCTION MATERIAL

#### Introduction.

The world around us consists of innumerable substances. The number of elements, natural as well as man-made, is over a hundred. Naturally, a study of the properties of such a large number of substances becomes difficult. Chemists, therefore, have tried to group together the elements having similar properties. They believe that such a classification would make the study of elements as well as their compounds easy and profitable. Such a grouping has been made and attempts are on to improve the grouping to make it more systematic. In chapter 7, a study of the properties of the noble gases, alkali metals and halogens has been made. These properties clearly indicate that separate groups could be formed for the noble gases, the alkali metals and the halogens. In this chapter, the student is introduced to the general principles of this grouping.

#### Outline.

1. The periodic law and the periodic table (8.1)
2. Elements of the second row (8.2)
3. Elements of the third row (8.3)

### Concepts.

1. The properties of elements are a periodic function of their atomic numbers.
2. The periodicity in properties of the elements is a consequence of their electronic configuration.

### SCHEDULE

Lesson	Experiment	Postlab	Demonstration	Text covered
1.	—	Discussion	—	8.1
2.	—	Discussion	—	8.1
3.	—	—	Demonstration I	8.2
4.	—	Discussion	Demonstration II Demonstration III	8.2 8.3
5.	—	Discussion	—	8.3
6.	—	—	—	Review & quiz

### Development.

Display as many elements as possible on the table. The following elements are easily available : copper, iron, aluminium, sodium, carbon, sulphur, magnesium, bromine, mercury, silver, tin, gold, zinc and platinum. In addition to these, gases like hydrogen, nitrogen, oxygen and chlorine can be prepared and jars filled with them can be kept on the table. The students should be asked to suggest a way of grouping the elements. The grouping suggested by the students at this stage will most probably be according to their : (i) state, (ii) colour, (iii) lustre etc. At this stage point out the limitations in the groupings made by the students on the basis of a few observable properties. Then suggest another grouping based on metallic and non-metallic nature of the elements with which the students are already familiar. At the end point out the limitations of this grouping as well.

At this stage, give a very brief historical account of the earlier attempts of classification to the students. A historical account creates interest in the minds of the students and also gives them an idea of this line of thinking in scientific development.

## 8.1 The periodic law and the periodic table

Explain the concept of periodicity by giving simple illustrations such as periodic recurrence of the seasons, the rising and setting of the sun and the swinging motion of the pendulum.

Then proceed by arranging the element in a horizontal row in the order of increasing atomic numbers and point out that the properties go on changing gradually from lithium to neon. Hydrogen, the element with atomic number 1, has some resemblance to lithium. Neon (At. No. 10) resembles helium (At. No. 2) This fact can explain the position of hydrogen and helium in the periodic table. Thus the first period contains only two elements. (The peculiar position of hydrogen need not be discussed in detail). Sodium, the element next to neon, has properties similar to those of lithium and finds its appropriate place below lithium. Bring out clearly the idea of periodicity in the properties of elements by discussing the arrangement of hydrogen, lithium and sodium

The arrangement of the elements up to chlorine should be discussed to emphasise the concept that periodicity of properties is not restricted to the group of elements, lithium and sodium or helium and neon. The properties of every element in the second period are found to be similar to those of the corresponding eighth element in the third period.

The discussion in this section lead to the concept that the properties of elements are a periodic function of their atomic numbers. This is understandable because the atomic number of an element is the index of its electronic configuration. Thus, the atomic number is a fundamental property governing the periodicity of the properties of elements.

## 8.2 Elements of the second row

Demonstrations which deal with some properties of the oxides of the second and the third row elements should be taken up first. In the first demonstration a gradation in the increasing acidity of the oxides of B, C, N is shown. Oxides of other elements from the second (Li, Be, F) are not easily available. In fact it is difficult to prepare fluorine oxide.

## Demonstration I : Some properties of the oxides of second row elements

Time required : 20 minutes

### Materials required :

Test tubes	6	Boron (III) oxide (Boric oxide)	0.5 g
Test tube stand	1	Gas jar filled with CO <sub>2</sub> gas	
Gas jars	2	Gas jar filled with NO <sub>2</sub> gas	
		Distilled water	30 ml
		Universal indicator	

Carbon dioxide gas may be prepared in a Kipp's apparatus by the action of dil. HCl on marble pieces. Collect the gas CO<sub>2</sub> by the upward displacement of air. Nitrogen dioxide should be prepared by the action of conc. nitric acid (25 ml) on copper turnings (5 g). Each of these gases should be collected in 2 or 3 jars. If boric oxide is not available, boric acid (available in medical shops) may be used.

### Procedure

- 1) Add about 0.5 g of boric oxide to 10 ml of distilled water in a test tube. Boil the solution gently for a minute. Add 2 or 3 drops of universal indicator to the solution and keep the test tube in a test tube stand.  
The solution acquires an yellow colour. pH about 6.
- 2) Add 10 ml of distilled water to the gas jar containing CO<sub>2</sub> gas. Cover the jar with a glass plate. Shake the jar and transfer the solution to a test tube. Add 2 or 3 drops of universal indicator to it and keep the tube in the stand.

The solution acquires orange colour. pH about 4.

- 3) Add 10 ml of distilled water to the gas jar containing  $\text{NO}_2$ . Cover the jar with a glass plate, shake it and transfer the solution to a test tube. (Note : If the solution becomes yellowish, warm the solution a little to remove dissolved  $\text{NO}_2$ ). Then add to the colourless solution the same number of drops of universal indicator (as in the earlier cases) and keep the tube in the stand.

The solution acquires a red colour. pH about 2.

Compare the colours of the solutions in the three test tubes with the universal indicator chart and show that the acidity increases from  $\text{B}_2\text{O}_3$  to  $\text{CO}_2$  to  $\text{NO}_2$ . The discussion of the observations may be done with the help of Table 8.1 in the text,

### **Demonstration II : Some properties of the elements of the third period**

Time required : 30 minutes

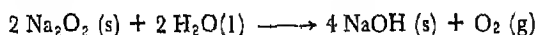
#### **Materials required**

Test tubes 6	Sodium peroxide	1 g
Beaker 100 ml 1	Magnesium ribbon piece	5 cm
Test tube holder	Sulphur	100 mg
Graduated cylinder 10 ml 1	Phosphorus	100 mg
Wooden splinters	Iron (II) sulphide	1 g
Deflagrating spoon	Sodium chloride	1 g
Tongs	Conc. sulphuric acid	4 ml
	200 M sodium hydroxide solution	4 ml
Gas jars with covers 2	(80 g solid in 1 litre of solution	
	Ammonium chloride	1 g
	Universal indicators of pH papers*	

(\* if not available, red and blue litmus papers may be used).

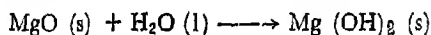
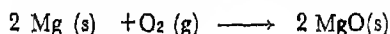
## Procedure

- (1) Add 5 ml of water to 1 g of sodium peroxide. (Caution : Sodium peroxide being a dangerous chemical should not be touched with fingers. Use a spatula) Use a glowing splinter to identify the gas produced. Add a piece of the indicator paper (or a red litmus paper and one of blue) to the contents of the test tube. Let the students note the result. Glowing splinter bursts into flame. Red litmus turns blue. Blue litmus remains unchanged.

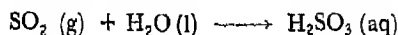
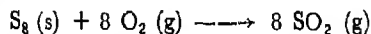


- (2) Hold one end of a magnesium ribbon with a pair of tongs and ignite the other end in a Bunsen burner flame. Hold the burning magnesium over a beaker so that the oxide produced drops into the beaker. Add 5 ml of water to the product and heat gently. Test the liquid with red and blue litmus papers.

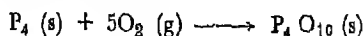
Red litmus turns blue. Blue litmus remains unchanged.



- (3) Put a pea-size of sulphur powder in a deflagrating spoon, ignite it in the burner flame and lower it into the gas jar and cover it with a glass plate. Add 10 ml of water to the gas jar contents, cover it with a glass plate and shake the contents. Transfer the contents to a test tube and test with litmus papers. Blue litmus turns red. Red litmus remains unchanged.



- (4) Repeat 3 using a pinch of red phosphorus. Blue litmus turns red. Red litmus remains unchanged



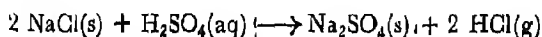
Let the students record in a tabular form the observations made in these reactions.



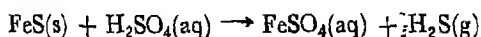
Experiment	Observation
<p>1. To 1 g of <math>\text{Na}_2\text{O}_2</math> add 5 ml of <math>\text{H}_2\text{O}</math></p> <p>Expose a glowing splinter to the gas.</p> <p>Add a piece of red litmus paper.</p> <p>Add a piece of blue litmus paper.</p>	<p>Colourless, odourless gas is produced briskly.</p> <p>Glowing splinter bursts with flame.</p> <p>Red litmus paper turns blue.</p> <p>Blue litmus paper remains unchanged.</p>
<p>2. Hold a magnesium ribbon with tongs, burn in a flame. Collect the white substance in a beaker and add water. Shake well. Test the resulting liquid with</p> <p>(a) red litmus paper</p> <p>(b) blue litmus paper</p>	<p>Magnesium ribbon burns with dazzling brilliance forming a white substance.</p> <p>(a) Red litmus paper turns blue.</p> <p>(b) Remains unchanged</p>
<p>3. Burn a pea-size sulphur in a deflagrating spoon in a jar. Cover it with a glass plate. Add 10 ml of water to it. Shake well. Test the resulting solution with</p> <p>(a) blue litmus paper</p> <p>(b) red litmus paper</p>	<p>Sulphur burns with a blue flame. A colourless gas with a choking smell is produced.</p> <p>(a) Blue litmus paper turns red.</p> <p>(b) Remains unchanged</p>
<p>4. Repeat the above experiment using red phosphorus and test the resulting solution with</p> <p>(a) Blue litmus paper</p> <p>(b) Red litmus paper</p>	<p>Copious white fumes are formed</p> <p>(a) Blue litmus paper turns red</p> <p>(b) Remains unchanged</p>

The following three steps show the reactions of typical hydrides.

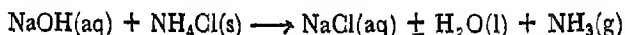
- (5) Add 2 ml of concentrated sulphuric acid to 1 g of sodium chloride and heat it gently. Test the gas evolved (HCl) with moistened litmus papers. White fumes are evolved. Red litmus remains unchanged. Blue litmus turns red.



- (6) Repeat step 5 using 1 g iron (II) sulphide with 2 ml of dilute sulphuric acid. Test the gas with litmus and identify the gas with a lead acetate paper. Colourless gas with an offensive smell is liberated. Red litmus remains unchanged. Blue litmus turns red. The lead acetate paper turns black.



- (7) Add 2 ml of 2.5 M sodium hydroxide solution to 1 g of ammonium chloride. Heat gently and test the gas produced with moist red and blue litmus papers and with turmeric paper, (Note : Since it is not feasible to do a test with phosphine, the hydride of phosphorus, ammonia is chosen) Colourless gas with pungent odour is produced. Moist red litmus paper turns blue. Blue litmus paper remains unchanged. Turmeric paper turns bluish-black or greenish black.



The students may be asked to write down equations for the reactions evolved. They may also be asked to write down the ionic equations showing the reactions of hydrides in steps 5-7 with water.

Ionic equations showing the reactions of hydrides (formed in steps 5-7) with water are as follows :

- i)  $\text{HCl(g)} + \text{H}_2\text{O(l)} \longrightarrow \text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq})$
- ii)  $\text{H}_2\text{S(g)} + \text{H}_2\text{O(l)} \longrightarrow \text{HS}^-(\text{aq}) + \text{H}^+(\text{aq})$   
 $\text{HS}^-(\text{aq}) + \text{H}_2\text{O} \longrightarrow \text{S}^{--}(\text{aq}) + \text{H}_3\text{O}^+(\text{aq})$
- iii)  $\text{NH}_3(\text{g}) + \text{H}_2\text{O(l)} \longrightarrow \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$

### Demonstration III : Properties of third row elements

Time required : 30 minutes

#### Materials required

Test tubes	: 12	0.5 M NaOH Solution	10ml
Test tube stands	: 2	(20 g solid per litre of solution)	
Test tube holder	: 1	Magnesium oxide	0.2 g
		Silica gel ( $\text{SiO}_2, n\text{H}_2\text{O}$ )	0.2 g
		Phosphorus pentoxide	0.1 g
		0.5 M Sulphuric acid	10 ml
		(14 ml diluted to 1 litre)	
		0.5 M perchloric acid	10 ml
		(14 ml diluted to 1 litre)	
		Universal indicator	

#### Procedure

Prepare a set of six labelled test tubes containing the solution as given below :

1. 10 ml of NaOH solution.
2. 0.2 g of magnesium oxide in 10 ml of water.
3. 0.2 g of silica gel in 10 ml of water.
4. About 0.1 g of phosphorus pentoxide in 10 ml of water  
(Caution : Reaction is very vigorous)
5. 10 ml 0.5 M sulphuric acid solution (solution of  $\text{SO}_3$  in water)
6. 10 ml 0.5 M perchloric acid solution (solution of  $\text{Cl}_2\text{O}_7$  in water)

Add equal number of drops of universal indicator to these solutions. Let the students observe the colours of the solutions. These observations will give an idea of the graded change from alkaline to acidic character of the solution, as one passes from sodium hydroxide (NaOH) to perchloric acid ( $\text{HClO}_4$ ). The colours should be compared with the universal indicator chart.

**Note :** It is practically impossible to weigh exactly 0.1 g  $\text{P}_2\text{O}_5$  as it is highly hygroscopic. A pinch of  $\text{P}_2\text{O}_5$  can be taken with a spatula and added to 10 ml water. If  $\text{P}_2\text{O}_5$  is not available, phosphoric acid ( $\text{H}_3\text{PO}_4$ ) which is a solution of  $\text{P}_2\text{O}_5$  in water can be used.

It should be made clear to the students that the solutions in the six test tubes are in fact the solutions of corresponding oxides of the third row elements. They should be asked to note down the observations and compare with the universal indicator chart in a tabular form as given below :

No.	Oxide in solution	Colour of the solution with the universal indicator	Acidity or alkalinity in pH units
1.	$\text{Na}_2\text{O}$	Purple	12
2	$\text{MgO}$	Blue	8
3.	$\text{SiO}_2$	Yellow	6
4.	$\text{P}_4\text{O}_{10}$	Orange	4
5.	$\text{SO}_3$	Red	2
6.	$\text{Cl}_2\text{O}_7$	Intense red	0

The discussion on observations made in demonstrations II and III should be made to bring out the variation in properties of the third row elements. The properties given in Table 8.2 may also be discussed to bring out the variation of properties.

### 8.3 Elements of the third row

The main idea to be brought out in this section is the variation in the reactivity of the elements in a group. This can be done by recalling the action of water on sodium and potassium (demonstration IV, chapter 7). It can be shown that under controlled conditions potassium reacts with water more vigorously than sodium, rubidium reacts more vigorously than potassium and so on. Thus lithium is the least reactive metal and cesium, the most. This order of reactivity is true for all reactions.

Refer to the properties of halogens that the student has already studied. Also recall the displacement reactions of halogens (Experiment 18 and demonstration III chapter 7). The discussion would lead to the conclusion that in the case of halogens the reactivity decreases from fluorine to iodine.

Discussions on this section made in the light of the material given in the text should lead to the concept that the periodicity of the properties of elements is closely related to their electronic configuration.

## BACKGROUND INFORMATION

### Early attempts at classification - Historical development

In the early classification, elements were divided into two broad groups, metals and nonmetals. Metals were characterised by high densities, good conduction of heat and electricity, and a typical shining appearance, called metallic lustre. The non metals on the other have low densities, are poor or non-conductors of heat and electricity and have no shining appearance. Some elements which exhibited properties of metals as well as non-metals were classed as 'metalloids'.

#### 8.1 The periodic law and the periodic table

The first significant step towards a systematic classification of elements was in 1817 by Dobereiner, who grouped the elements with similar properties into groups of three, such as Li-Na-K, Ca-Sr-Ba and Cl-Br-I. He called it as the "Law of triads". He showed that the atomic weights of these elements are in increasing order and that the atomic weight of the middle element in a triad is an average of the atomic weights of the other two elements. Thus the atomic weights were used in the arrangement of elements for the first time. It was not possible however to show the existence of such triads in the case of a majority of other elements.

Newlands, a British chemist, arranged the elements in the order of increasing atomic weights, starting from hydrogen. He showed that "the eighth element", starting from a given one, is a kind of a repetition like the eighth note in an 'octave of music'. In other words, the properties repeat after a definite period. His arrangement is shown below :

H						
	Li	Be	B	C	N	O
F	Na	Mg	Al	Si	P	S
Cl	K	Ca	Sr	Tl	Mn	Fe
Ni	Cu	Zn	Y	In	As	Se
Br						

It was clearly noted that Newlands' law of octaves was true only for the first 17 elements except H (for the elements shown in between the dotted lines).

Other chemists were also busy with the same problem. In 1864 Lothar Meyer in Germany determined the atomic volumes, i. e. the volumes of one gram-atom of many elements in  $\text{cm}^3$ , in their solid state. He plotted atomic volumes of elements against their atomic weights. All the alkali metals Li, Na, K, Cs, Rb having similar properties lie on the peaks of the atomic volume curve, while the halogen elements occupy another set of similar positions. Lothar Meyer's curve shows that the atomic volume of elements is a periodic property.

In 1869 Mendeleef, a Russian chemist, put forth his idea of a periodic table obtained on the basis of atomic weights. He stated that the properties of elements and their compounds are periodic function of their atomic weights. He arranged the elements in horizontal rows (periods) and vertical columns (groups). In two cases (Te—I and Co—Ni), Mendeleef placed the elements in a reverse order of their atomic weights. He pleaded for these reversals on the basis of their properties. There are other triumphant features of Mendeleef's table, like the following.

- (1) Mendeleef could settle some of the controversial atomic weights from the positions of the elements in the periodic table. For instance, the atomic weight of beryllium as known from the experiments could be either 9 or 13.5. *It could not be known with certainty.* When Mendeleef placed it, on the basis of its properties, between Li (atomic weight 7) and B (atomic weight 11), the atomic weight of Be could be taken only as 9.
- (2) In addition, Mendeleef could as well predict the existence of some more elements from the gaps he deliberately left in his table; e.g. the element with atomic weight 72 was missing which he called ekasilicon, suggesting that an element similar to silicon, though then unknown, must exist. Further he predicted its properties. The element was actually discovered in 1888 and was given the name Germanium (Ge). The following table shows the correctness of Mendeleef's predictions.

<b>Eka-silicon (predicted) (1869)</b>	<b>Germanium (discovered) (1888)</b>
1. Atomic weight 72	Atomic weight 72.6
2. Density 5.5 g/cc	Density 5.47 g/cc
3. Dull grey metal	Greyish white metal
4. Will form a white powder of $\text{EsO}_2$ by heating in oxygen	Forms white oxide $\text{GeO}_2$ on heating in oxygen
5. No action of alkali	No action of alkali
6. Very little action of acids	No action of $\text{HCl}$ , dissolves in aqua regia
7. Will form a tetravalent chloride with B. P. less than $100^\circ\text{C}$	Forms $\text{GeCl}_4$ , B. P. $86^\circ\text{C}$

The main success of his table was in the prediction of more such elements as Eka-boron (scandium), Eka-aluminium (Gallium) and Eka-silicon (Germanium). The periodic table which is based on atomic numbers is an extension of Mendeleef's original table.

After the discovery of the radioactive nature of some of the elements in 1896 by Becquerel, a new era started in the development of the ideas of the structure of the atom. Rutherford was able to find out the number of unit positive charges,  $i.e.$ , number of protons in the nucleus of atoms of some elements. An epoch-making discovery was made by Moseley in 1913, while studying the X-ray spectra of elements. He found that the number of positive charges on the atom of an element is equal to the serial number (ordinal number) of the element in Mendeleef's periodic table. This number of positive charges was named by him as the atomic number. Moreover the reserved positions of Co-Ni, Te-I (and of Ar-K observed later) as predicted by Mendeleef were found to be correct if the elements were arranged in the order of increasing atomic numbers instead of atomic weight.

Scientists have been trying to construct better forms of the periodic table on the basis of atomic number as a fundamental property. The long form, the extended form, the spiral form, the circular form, the keyhole form and the laminar form are all the results of such attempts. These attempts are likely to go on. Even though these different forms are known, the long form is more popularly used as a teaching aid, as it reveals best the electron configuration of the elements, and emphasizes the order of filling of the electron shells in relation to the order of increasing energy states. This aspect of the modern approach is more clearly brought out in the table proposed by Longuet-Higgins. The very fact that so many tables have been proposed points out the variety of opinions that exist and indicates at the same time how difficult it would be to prescribe the 'best' periodic table.

It is evident from the discussion of the periodic table that those properties which depend upon the electron configuration of an atom will vary periodically with atomic number. Some of the more common periodic properties are listed below.

### **Properties dependent upon the electronic configuration**

Atomic radius (and volume)	Density
Ionic radius	Melting point
Ionization potential	Boiling point
Electron affinity	Heats of fusion, vapourisation and sublimation
Electronegativity	Heats of solvation of ions

Valency	Heat of formation for a given type of compound
Standard Redox potential	Hardness, malleability
Magnetic behaviour	Compressibility
Thermal and electrical conductivities	Coefficient of expansion
Ionic mobility	Parachor
Optical spectrum	Refractive index

Not only are these properties periodic functions of the atomic weights of the elements but they are also periodic for the analogous compounds of the elements.

## 8.2 Elements of the second and the third periods

The earlier chapter (7) has focussed on the three well-defined groups of the periodic table, the alkali metals, the halogens and the noble gases. How about the other vertical groups of the table? Rather than study each group in a manner similar to that used for the three groups studied, the second and third period elements are chosen as representatives of their particular groups. The emphasis is placed on changes in properties and especially on their gradation. *The pattern in the formation of oxides of these elements is evident in tables 8.1 and 8.2. This sort of pattern is also observed in the formation of chlorides and hydrides.*

## 8.3 Properties of elements in a group

The original periodic tables were based upon the observed similarities in the properties of certain groups of elements: The members of the alkali metals group, for instance, were listed in a single column because it was observed that the members of this family had strikingly similar properties. An understanding of the cause of this similarity in properties, however, had to wait for a satisfactory model of the atom. In terms of our present model, the electronic configurations of the alkali metals (Li, Na, K, Rb and Cs) show that each member of the family has one electron in the outermost shell, with a similar internal electron structure. It is this similarity in electron configuration that leads to the similarities in chemical and physical properties of the periodic table. There is also a more or less regular change in those similar properties in going from the lightest to the heaviest member of the family. This is observed in other families also.



## Quiz

Each question is followed by four answers. Only one of them is correct. Select the correct one and indicate it in the brackets provided.

1. The periodic law states that the properties of elements are a periodic function of their ( )  
(a) atomic weight (b) mass number  
(c) atomic number (d) density.
2. In the periodic table as we go across a period of the elements from left to right the metallic character of the elements ( )  
(a) increases (b) decreases  
(c) does not change (d) changes irregularly.
3. Magnesium oxide is ( )  
(a) acidic (b) basic  
(c) neutral (d) amphoteric.
4. The halogens are placed in a group because ( )  
(a) they have the same electronic configuration  
(b) they have the same number of electrons  
(c) they are all very reactive  
(d) they have similar electronic configuration.
5. The acidic character of the oxides of elements in a row ( )  
(a) increases (b) decreases  
(c) remains constant (d) changes irregularly.
6. The atomic number of an alkali metal is less than that of the following halogen element by ( )  
(a) two (b) four  
(c) six (d) eight.
7. The element which forms a dipositive ion is ( )  
(a) sodium (b) fluorine  
(c) calcium (d) sulphur.



## QUESTIONNAIRE

### On the school chemistry laboratory manual, text book and teacher's guide for class 8

*N. B. —In case the space provided against each question is insufficient, please attach additional papers. Specific and concrete suggestions for any modification, alteration and improvement of the material will be appreciated.*

1. Are the experiments in the laboratory manual simple enough to be easily done by the pupils?

If not, which experiments, do you feel are difficult?

How would you like them to be modified?

2. Are the experiments suited to arouse the curiosity and interest of the pupils?

Do they lead to an appreciation of the basic principles of chemistry?

If not, what improvements do you suggest?

3. In general, are the experiments in the laboratory manuals well planned?

4. Do you think that the experiments mentioned in the manuals are adequate?

If not, what additional experiments would you like to be introduced and at what places?

5. Do you think that the materials required for the experiments would be easily available ?

If not, can you suggest alternatives using more easily available materials to bring out the same ideas ?

6. What, in your opinion, is the scope and extent of improvisation that can and may be made to replace the equipment mentioned ?

7. What, in the context of the present condition of our schools, is your opinion about the feasibility of the idea of pupil experimentation ?

8. The material in the text has been prepared; to a large extent, to help the pupils learn chemistry by drawing conclusions from his own observations in the laboratory.

Do you think that this objective has been broadly realised ?

9. Is the language on the whole simple enough for an easy understanding of the subject ?

If not, point out where the language could be simplified.

10. Do you think that the material as developed in the text makes for continued and logical development of ideas and concepts ?

11. Are the contents of each unit within the comprehension of pupil of the age group of 12+ and 14+ ?

If not, is it possible to express the same concept in a simpler way ?

12. Do you think that the sequence of the units is satisfactory? If not, please suggest how the sequence may be rearranged to enable a better development of the subject

Give reasons for your rearrangement.

13. Are the diagrams and illustrations suggestive and self-explanatory?

State how they may be improved.

Do you think that the present illustrations are adequate?

If not, which additional diagrams do you suggest?

14. Do the demonstration experiments mentioned in the teacher's guide bring about the concepts?

If not, what alternative experiments would you suggest?

15. The teacher's guide is intended to enable the teacher to teach the subject confidently and more effectively. It guides him to plan the lesson and provides background information to extend the sphere of his knowledge and answer questions which might be put by intelligent pupils. How far does this teacher's guide serve the purpose for which it is intended?

16. Do you think that the teacher's guide contains sufficient information for the teacher to present the topic effectively?

If not, what additional information may be given and where?

17. The aim of this effort is to introduce the pupils to modern concepts in chemistry and promote a sense of experimentation and inquiry in him. Do you think that this would be achieved by the text and laboratory work presented?

What alterations would you suggest in this regard?

18. What other pertinent comments and suggestions would you like to make?

Signature

Name :

Station :

Address :

Date :

